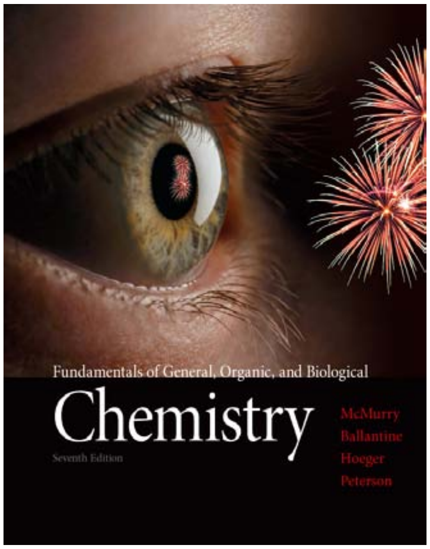
Chapter 2 Lecture



Fundamentals of General, Organic, and Biological Chemistry

7th Edition

McMurry, Ballantine, Hoeger, Peterson

Chapter Two

Atoms and the Periodic Table

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ALWAYS LEARNING PEARSON

Outline

2.1 Atomic Theory
2.2 Elements and Atomic Number
2.3 Isotopes and Atomic Weight
2.4 The Periodic Table
2.5 Some Characteristics of Different Groups
2.6 Electronic Structure of Atoms
2.7 Electron Configurations
2.8 Electron Configurations and the Periodic Table
2.9 Electron-Dot Symbols

Goals

1. **What is the modern theory of atomic structure?**
   Be able to explain the major assumptions of atomic theory.
2. **How do atoms of different elements differ?**
   Be able to explain the composition of different atoms according to the number of protons, neutrons, and electrons they contain.
3. **What are isotopes, and what is atomic weight?**
   Be able to explain what isotopes are and how they affect an element's atomic weight.

Goals, *Continued*

4. **How is the periodic table arranged?**
   Be able to describe how elements are arranged in the periodic table, name the subdivisions of the periodic table, and relate the position of an element in the periodic table to its electronic structure.
5. **How are electrons arranged in atoms?**
   Be able to explain how electrons are distributed in shells and subshells around the nucleus of an atom, how valence electrons can be represented as electron-dot symbols, and how the electron configurations can help explain the chemical properties of the elements.

2.1 Atomic Theory

Questions about the fundamental nature of matter can be traced as far back as the Greek philosophers.

- Aristotle believed that matter could be divided indefinitely.
- Democritus argued that there was a limit.

The smallest particle that an element can be divided into and still be identifiable is an **atom**, from the Greek *atomos*, meaning "indivisible."

2.1 Atomic Theory

Chemistry is founded on four assumptions which together make up **Atomic Theory**.

1. All matter is composed of atoms.
2. The atoms of a given element differ from the atoms of all other elements.
3. Chemical compounds consist of atoms combined in specific ratios. Only whole atoms can combine.
4. Chemical reactions change only the way atoms are combined in compounds.

2.1 Atomic Theory

Atoms are composed of tiny **subatomic particles** called *protons, neutrons*, and *electrons*.

- **Protons** carry a positive electrical charge.
- **Neutrons** have a mass similar to that of a proton, but are electrically neutral.
- **Electrons** have a mass that is only 1/1836 that of a proton and carry a negative electrical charge.
  - Electrons are so much lighter than protons and neutrons that their mass is usually ignored.

2.1 Atomic Theory

TABLE 2.1 A Comparison of Subatomic Particles

		Mass		
Name	**Symbol**	**(Grams)**	**(amu)**	**Charge (Charge Units)**
Proton	p	$1.672\,622 \times 10^{-24}$	1.007 276	+1
Neutron	n	$1.674\,927 \times 10^{-24}$	1.008 665	0
Electron	e^-	$9.109\,328 \times 10^{-28}$	$5.485\,799 \times 10^{-4}$	−1

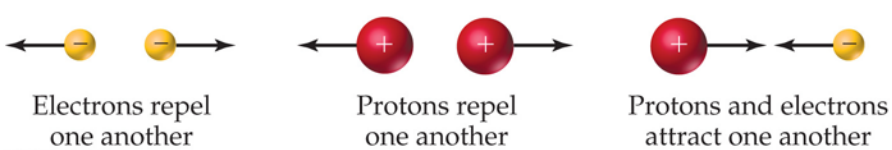
- The masses of atoms and subatomic particles are expressed on a relative mass scale.
- The basis for the scale is an atom of carbon that contains 6 protons and 6 neutrons. This carbon atom is assigned a mass of exactly 12 **atomic mass units** (**amu**) or *daltons*.

2.1 Atomic Theory

- The protons and neutrons are packed closely together in a dense core called the **nucleus**. Surrounding the nucleus, the electrons move about rapidly through a large volume of space.
- The relative size of a nucleus in an atom is the same as that of a pea in the middle of this stadium.

2.1 Atomic Theory

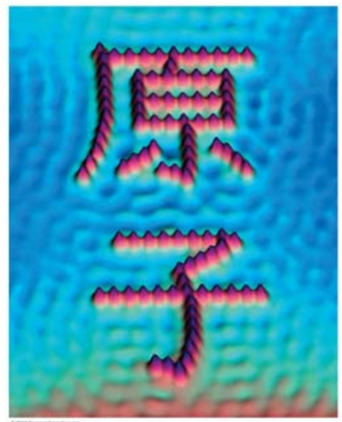
- The structure of the atom is determined by interplay of different forces.
- Opposite electrical charges attract each other, like charges repel each other.
- Protons and neutrons in the nucleus are held together by the *nuclear strong force*.



2.1 Atomic Theory

Are Atoms Real?

- John Dalton proposed atomic theory in 1808.
- J. J. Thomson determined the mass of the electron.
- Robert Millikan identified the charge on the electron.
- Ernest Rutherford deduced the existence of the nucleus.
- The *scanning tunneling microscope* makes individual atoms visible.

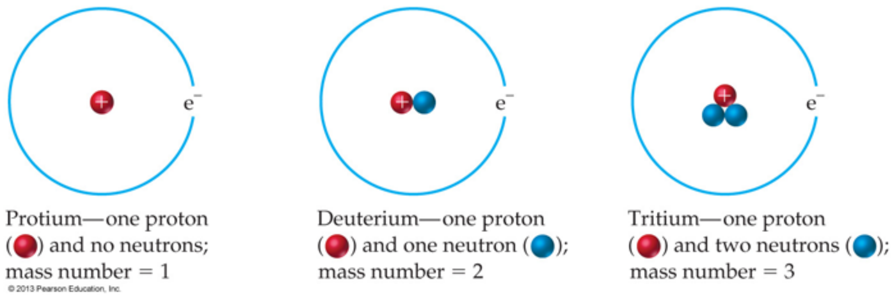


2.2 Elements and Atomic Number

- **Atomic Number (*Z*)** is the number of protons in atoms of a given element. All atoms of a particular element have the same number of protons in the nucleus.
- Atoms are neutral because the number of positively charged protons and the number of negatively charged electrons are the same in each atom.
- **Mass Number (*A*)** is the sum of the protons and neutrons in an atom.

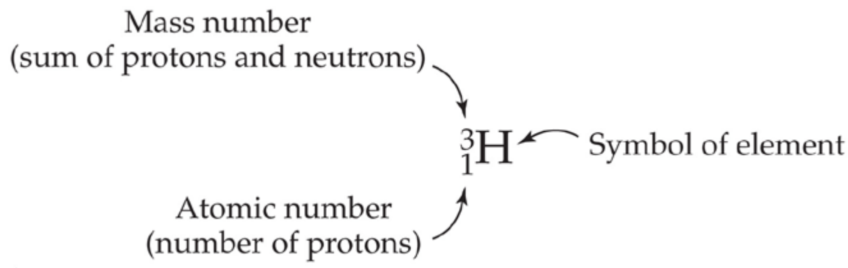
2.2 Elements and Atomic Number

- **Isotopes** are atoms with identical atomic numbers but different mass numbers.
- Protium, deuterium, and tritium are three isotopes of hydrogen.



2.2 Elements and Atomic Number

- A specific isotope is represented by showing its mass number (*A*) as a superscript and its atomic number (*Z*) as a subscript in front of the atomic symbol.
- For example, the symbol for tritium is:



2.2 Elements and Atomic Number

- The isotopes of most elements do not have distinctive names.
- The mass number (*A*) is given after the name of the element.
- The isotope used in nuclear reactors is usually referred to as uranium-235, or U-235
- Most naturally-occurring elements are mixtures of isotopes.

2.2 Elements and Atomic Number

Atomic Weight is the weighted average mass of an element's atoms.

- The individual masses of the naturally occurring isotopes and the percentage of each must be known.
- The atomic weight is calculated as the sum of the masses of the individual isotopes for that element.

$$\text{Atomic weight} = \Sigma[(\text{isotope abundance}) \times (\text{isotope mass})]$$

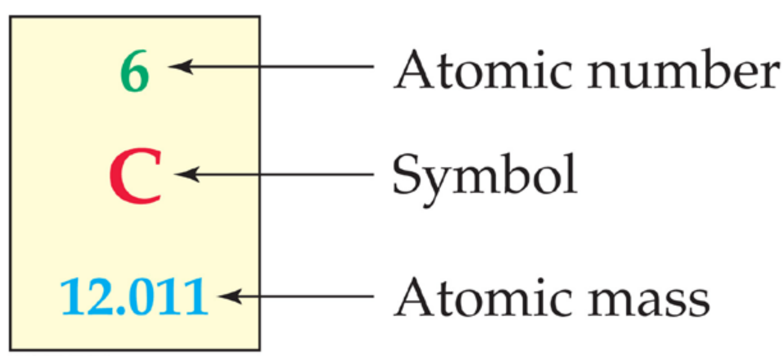
- The Greek symbol Σ indicates summing of terms.

2.4 The Periodic Table

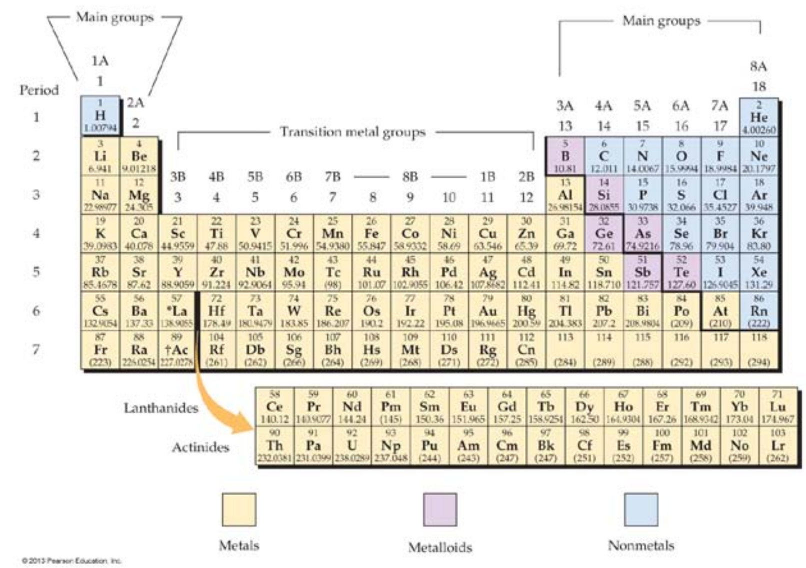
- Only 10 elements have been known since the beginning of recorded history.
- In the late 1700s and early 1800s, as more elements were discovered, chemists began to look for similarities among elements.
- Johann Döbereiner observed in 1829 that there were several *triads* that appeared to have similar chemical and physical properties.
- Dmitri Mendeleev organized elements by mass, then into columns based on chemical behavior.

2.4 The Periodic Table

Mendeleev produced the forerunner of the modern periodic table, which has boxes for each element with the symbol, atomic number, and atomic weight.



2.4 The Periodic Table



2.4 The Periodic Table

The vertical **groups** on the periodic table have similar chemical properties and are divided into categories.

- **Main Groups—**The two groups on the far left (1–2) and the six on the far right (13–18) are the main groups.
- **Transition Metal Groups—**Elements in these groups are numbered 3 through 12.
- **Inner Transition Metals—**The 14 elements following lanthanum are the *lanthanides* and the 14 elements following actinium are the *actinides*. Together these are the inner transition metals.

2.4 The Periodic Table

- Elements (except hydrogen) on the left-hand side of the black zigzag line running from boron (B) to tellurium (Te) are *metals.*
- Elements to the right of the line are *nonmetals.*
- Most elements abutting the line are *metalloids.*
- Elements in a given group have similar chemical properties. Chlorine, bromine, iodine, and other elements in group 7A behave similarly.

2.4 The Periodic Table

All seven periods do not contain the same number of elements.

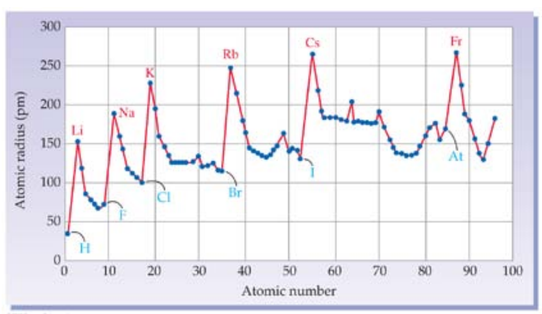
- The first period contains only 2 elements.
- The second and third periods each contain 8 elements.
- The fourth and fifth periods each contain 18 elements.
- The sixth and seventh periods each contain 32 elements.
- The 14 lanthanides and the 14 actinides are pulled out and shown below the others.

2.4 The Periodic Table

Groups are numbered in two ways, both shown in Figure 2.2.

- The **main group elements** are numbered 1A through 8A.
- The **transition metal elements** are numbered 1B through 8B.
- Alternatively, all 18 groups are numbered sequentially from 1 to 18.
- The **inner transition metal elements** are not numbered.

2.5 Some Characteristics of Different Groups



A graph of atomic radius versus atomic number shows a periodic rise-and-fall pattern. The maxima occur for group 1A elements, and the minima occur for group 7A elements.

2.5 Some Characteristics of Different Groups

- The periodic table has the name it does because the elements in it show an obvious *periodicity*.
- Many physical and chemical properties exhibit periodic behavior.
- The various elements in a group show remarkable similarities in their chemical and physical properties.
- Although the resemblances are not as pronounced as they are within a single group, *neighboring* elements often behave similarly as well.

2.5 Some Characteristics of Different Groups

Group 1A—Alkali metals

- Lithium (Li), sodium (Na), potassium (K), rubidium (Rb), cesium (Cs), and francium (Fr)
- Shiny, soft metals with low melting points
- React with water to form products that are highly alkaline
- Because of their high reactivity, alkali metals are never found in nature in a pure state.

2.5 Some Characteristics of Different Groups

Group 2A—Alkaline earth metals

- Beryllium (Be), magnesium (Mg), calcium (Ca), strontium (Sr), barium (Ba), and radium (Ra)
- Lustrous, silvery metals
- Less reactive than their neighbors in group 1A
- Never found in nature in the pure state

2.5 Some Characteristics of Different Groups

Group 7A—Halogens

- Fluorine (F), chlorine (Cl), bromine (Br), iodine (I), and astatine (At)
- Colorful and corrosive nonmetals
- Found in nature only in combination with other elements, such as with sodium in table salt (sodium chloride, NaCl)
- The group name, *halogen*, is taken from the Greek word *hals*, meaning salt.

2.5 Some Characteristics of Different Groups

Group 8A—Noble gases

- Helium (He), neon (Ne), argon (Ar), krypton (Kr), xenon (Xe), and radon (Rn)
- Colorless gases
- Labeled the "noble" gases because of their lack of chemical reactivity
- Helium, neon, and argon don't combine with any other elements. Krypton and xenon combine with very few.

2.5 Some Characteristics of Different Groups

Origin of Chemical Elements

- Within a second after the "big bang," subatomic particles began to form.
- After 3 minutes, the temperature had dropped enough that helium nuclei began to form.
- After millions of years, the universe cooled enough that electrons could bind to protons and helium nuclei, forming hydrogen and helium atoms.
- As clouds of hydrogen and helium condensed into stars, hydrogen fusion into helium began, releasing heat and light.
- As stars burned out, their core densities and temperatures increased, allowing the formation of larger nuclei.
- Supernovae resulted in the synthesis of even heavier elements.

2.6 Electronic Structure of Atoms

- The properties of the elements are determined by the arrangement of electrons in their atoms.
- This arrangement is understood using the *quantum mechanical model* developed by Erwin Schrödinger.
  - Electrons have both particle-like and wave-like properties.
  - The behavior of electrons can be described using an equation called a wave function.
  - Electrons are not perfectly free to move. They are restricted to certain energy values, or *quantized*.

2.6 Electronic Structure of Atoms

- A ramp is *not* quantized because it changes height continuously. Stairs *are* quantized because they change height by a fixed amount.
- Energy values available to electrons in an atom change only in steps.
- Wave functions also provide an electron with an "address" within an atom, composed of *shell, subshell*, and *orbital*.

2.6 Electronic Structure of Atoms

- Electrons in an atom are grouped around the nucleus into **shells**.
- The farther a shell is from the nucleus, the larger it is, the more electrons it can hold, and the higher the energies of those electrons.
  - The first shell (closest to the nucleus) can hold two electrons.
  - The second shell can hold 8 electrons.
  - The third shell can hold 32 electrons.

2.6 Electronic Structure of Atoms

- Within the shells, electrons are further grouped into **subshells** of four different types, identified as *s, p, d*, and *f* in order of increasing energy.
- 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 the fourth has *s*, *p*, *d* and *f* subshells.
- The number of subshells is equal to the shell number.
- A specific subshell is symbolized by writing the number of the shell, followed by the letter for the subshell.

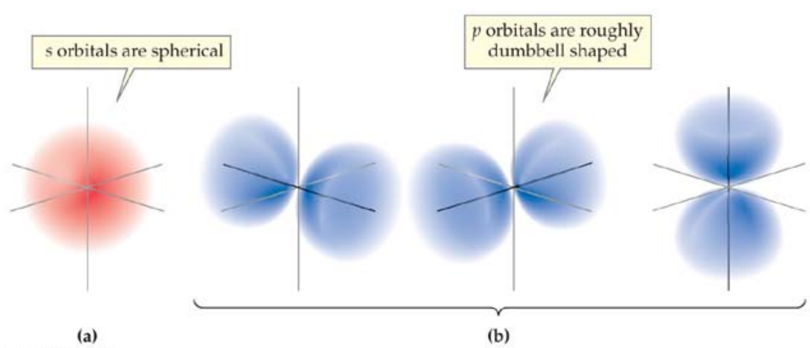
2.6 Electronic Structure of Atoms

- Within each subshell, electrons are grouped into **orbitals**, regions of space within an atom where the specific electrons are most likely to be found.
- An *s* subshell has 1 orbital, a *p* has 3 orbitals, a *d* has 5 orbitals, and an *f* has 7 orbitals.
- Each orbital holds two electrons which differ in a property known as *spin*.

Shell number:	1	2	3	4
Subshell designation:	*s*	*s* , *p*	*s* , *p* , *d*	*s* , *p* , *d* , *f*
Number of orbitals:	1	1 , 3	1 , 3 , 5	1 , 3 , 5 , 7

2.6 Electronic Structure of Atoms

Different orbitals have different shapes and orientations. Orbitals in *s* subshells are spherical, while orbitals in *p* subshells are roughly dumbbell-shaped.



2.6 Electronic Structure of Atoms

- Any orbital can hold a maximum of 2 electrons with opposite spin.
- The first shell has one 1*s* orbital and holds 2 electrons.
- The second shell holds 8 electrons; 2 in a 2*s* orbital and 6 in three 2*p* orbitals.
- The third shell holds 18 electrons; 2 in a 3*s* orbital; 6 in three 3*p* orbitals; and 10 in five 3*d* orbitals.
- The fourth shell holds 32 electrons; 2 in a 4*s* orbital; 6 in three 4*p* orbitals; 10 in five 4*d* orbitals; and 14 in seven 4*f* orbitals.

2.6 Electronic Structure of Atoms

The overall electron distribution within an atom is summarized in Table 2.2:

TABLE 2.2 Electron Distribution in Atoms

SHELL NUMBER:	1	2	3	4
Subshell designation:	*s*	*s* , *p*	*s* , *p* , *d*	*s* , *p* , *d* , *f*
Number of orbitals:	1	1 , 3	1 , 3 , 5	1 , 3 , 5 , 7
Number of electrons:	2	2 , 6	2 , 6 , 10	2 , 6 , 10 , 14
Total electron capacity:	2	8	18	32

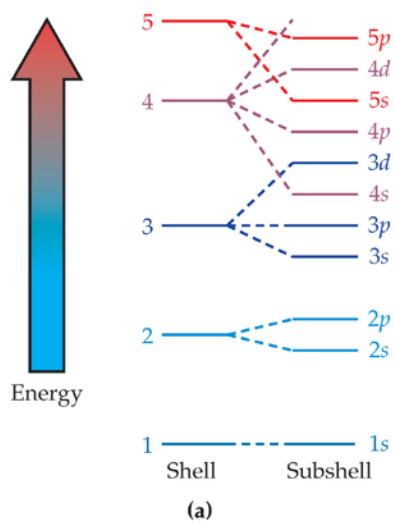
2.7 Electron Configurations

The exact arrangement of electrons in an atom's shells and subshells is the atom's **electron configuration**. It can be predicted by applying three rules.

Rule 1: Electrons occupy the lowest energy orbitals available. This is complicated by "crossover" of energies above the 3*p* level.

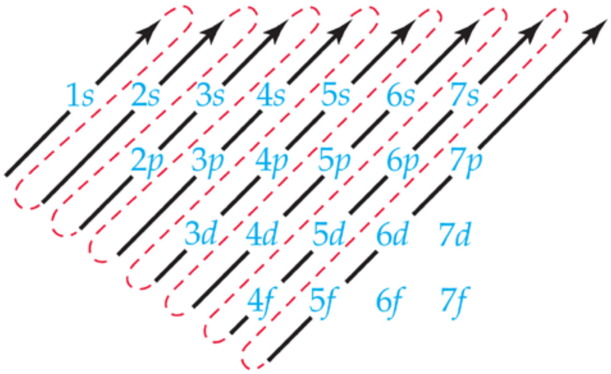
2.7 Electron Configurations

- Electrons fill orbitals from the lowest-energy orbitals upward.
- Above the 3*p* level, the energies of the orbitals in different shells begin to overlap.
- The energy level diagram can be used to predict the order in which orbitals are filled.



2.7 Electron Configurations

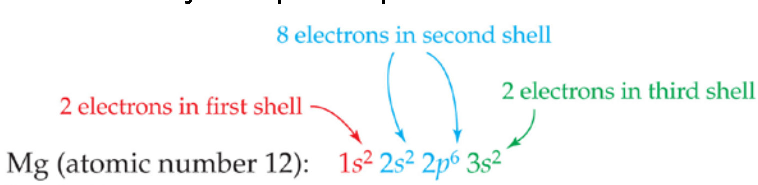
Below is a simple scheme to help remember the order in which the orbitals are filled.



(b)

2.7 Electron Configurations

- **Rule 2:** Each orbital can hold only two electrons, which must be of opposite spin.
- **Rule 3:** Two or more orbitals with the same energy are each half-filled by one electron before any one orbital is completely filled by addition of the second electron.
  - The number of electrons in each subshell is indicated by a superscript.



2.7 Electron Configurations

TABLE 2.3 Electron Configurations of the First 20 Elements

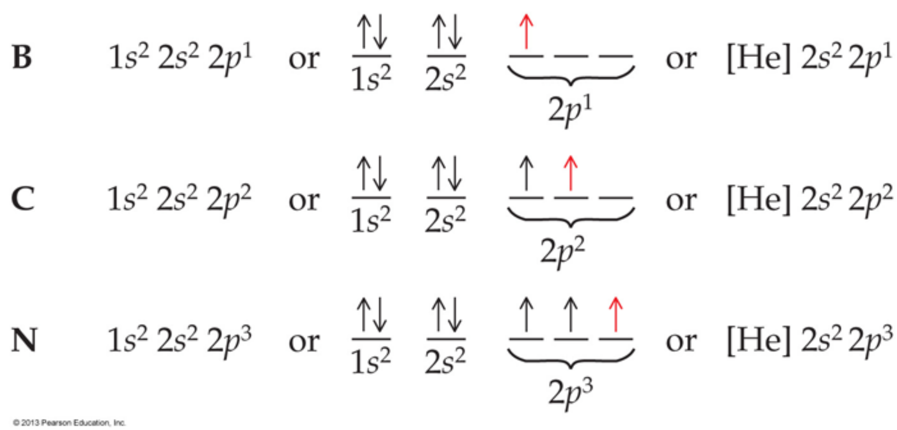
	Element	Atomic Number	Electron Configuration
H	Hydrogen	1	$1s^1$
He	Helium	2	$1s^2$
Li	Lithium	3	$1s^2 2s^1$
Be	Beryllium	4	$1s^2 2s^2$
B	Boron	5	$1s^2 2s^2 2p^1$
C	Carbon	6	$1s^2 2s^2 2p^2$
N	Nitrogen	7	$1s^2 2s^2 2p^3$
O	Oxygen	8	$1s^2 2s^2 2p^4$
F	Fluorine	9	$1s^2 2s^2 2p^5$
Ne	Neon	10	$1s^2 2s^2 2p^6$
Na	Sodium	11	$1s^2 2s^2 2p^6 3s^1$
Mg	Magnesium	12	$1s^2 2s^2 2p^6 3s^2$
Al	Aluminum	13	$1s^2 2s^2 2p^6 3s^2 3p^1$
Si	Silicon	14	$1s^2 2s^2 2p^6 3s^2 3p^2$
P	Phosphorus	15	$1s^2 2s^2 2p^6 3s^2 3p^3$
S	Sulfur	16	$1s^2 2s^2 2p^6 3s^2 3p^4$
Cl	Chlorine	17	$1s^2 2s^2 2p^6 3s^2 3p^5$
Ar	Argon	18	$1s^2 2s^2 2p^6 3s^2 3p^6$
K	Potassium	19	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$
Ca	Calcium	20	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$

2.7 Electron Configurations

- Electron configurations are described by writing the shell number and subshell letter in order of increasing energy. The number of electrons occupying each subshell is indicated by a superscript.
- A graphic representation can be made by indicating each orbital as a line and each electron as an arrow. The head of the arrow indicates the electron spin.
- A shorthand using noble gas configurations is very useful for large atoms.

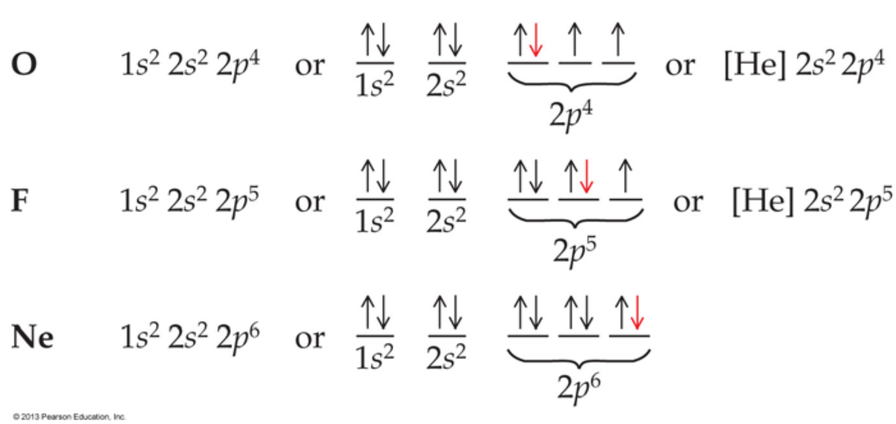
2.7 Electron Configurations

These are the electron configurations for B – N in which the $2p$ shell begins to fill.



2.7 Electron Configurations

These are the electron configurations for O – Ne in which the $2p$ shell is completed.



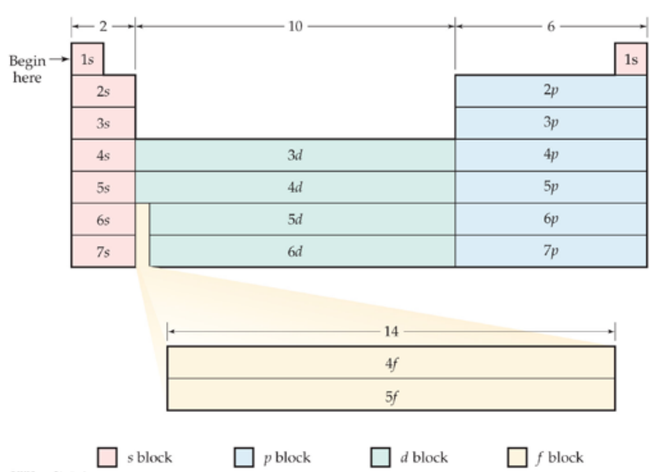
2.8 Electron Configurations and the Periodic Table

The periodic table can be divided into *blocks* of elements according to *the subshell filled last*.

- An ***s*-Block element** is a main group element that results from the filling of an *s* orbital.
- A ***p*-Block element** is a main group element that results from the filling of *p* orbitals.
- A ***d*-Block element** is a transition metal element that results from the filling of *d* orbitals.
- An ***f*-Block element** is an inner transition metal element that results from the filling of *f* orbitals.

2.8 Electron Configurations and the Periodic Table

The periodic table provides a method for remembering the order of orbital filling.



2.8 Electron Configurations and the Periodic Table

Elements in the same group of the periodic table have similar electron configurations in their valence shells.

- A **valence shell** is the outermost electron shell of an atom
- A **valence electron** is an electron in the valence shell of an atom.
- The group numbers from 1A through 8A give the numbers of valence electrons for the elements in each main group.
- Valence electrons are the most loosely held, and they are the most important in determining an element's properties.

2.8 Electron Configurations and the Periodic Table

Table 2.4 Valence-Shell Electron Configurations for Group 1A, 2A, 7A, and 8A Elements

Group	Element	Atomic Number	Valence-Shell Electron Configuration
1A	Li (lithium)	3	$2s^1$
	Na (sodium)	11	$3s^1$
	K (potassium)	19	$4s^1$
	Rb (rubidium)	37	$5s^1$
	Cs (cesium)	55	$6s^1$
2A	Be (beryllium)	4	$2s^2$
	Mg (magnesium)	12	$3s^2$
	Ca (calcium)	20	$4s^2$
	Sr (strontium)	38	$5s^2$
	Ba (barium)	56	$6s^2$
7A	F (fluorine)	9	$2s^2\ 2p^5$
	Cl (chlorine)	17	$3s^2\ 3p^5$
	Br (bromine)	35	$4s^2\ 4p^5$
	I (iodine)	53	$5s^2\ 5p^5$
8A	He (helium)	2	$1s^2$
	Ne (neon)	10	$2s^2\ 2p^6$
	Ar (argon)	18	$3s^2\ 3p^6$
	Kr (krypton)	36	$4s^2\ 4p^6$
	Xe (xenon)	54	$5s^2\ 5p^6$

2.9 Electron-Dot Symbols

Electron-dot symbol—An atomic symbol with dots placed around it to indicate the number of valence electrons.

TABLE 2.5 Electron-Dot Symbols for Some Main Group Elements

1A	2A	3A	4A	5A	6A	7A	NOBLE GASES
H·							He:
Li·	·Be·	·Ḃ·	·Ċ·	·Ṅ:	·Ö:	·F̤:	:N̤e:
Na·	·Mg·	·Ȧl·	·Ṡi·	·Ṗ:	·S̈:	·C̤l:	:Ä̤r:
K·	·Ca·	·Ġa·	·Ġe·	·Ȧs:	·S̈e:	·B̤r:	:K̤r:

2.9 Electron-Dot Symbols

Atoms and Light

- An atom with its electrons in their lowest-energy locations is said to be in its *ground state*.
- When electromagnetic energy collides with an atom, and the amount of electromagnetic energy is just right, an electron can be kicked up from its usual energy level to a higher one.
- With one of its electrons promoted to a higher energy, an atom is said to be *excited*.
- The excited state does not last long, because the electron quickly drops back to its ground state, releasing its extra energy in the process.

2.9 Electron-Dot Symbols

Atoms and Light

- Neon lights and fireworks are the result of this phenomenon.
- The concentration of metals in body fluids is measured by sensitive instruments relying on the same principle of electron excitation.

Chapter Summary

1. What is the modern theory of atomic structure?

- All matter is composed of *atoms*.
- An atom is the smallest and simplest unit which maintains the properties of the element.
- Atoms are made up of *protons, neutrons*, and *electrons*.
  - Protons have a positive electrical charge and are found in the *nucleus.*
  - Neutrons are electrically neutral, and are found in the nucleus.
  - Electrons have a negative charge, and are a relatively large distance from the nucleus. Most of the atom is empty space.

Chapter Summary, *Continued*

2. How do atoms of different elements differ?

- Elements differ according to the number of protons their atoms contain, a value called the *atomic number* (*Z*).
- All atoms of a given element have the same number of protons and an equal number of electrons.
- The number of neutrons in an atom is not predictable, but is generally as great or greater than the number of protons.
- The total number of protons plus neutrons in an atom is called the atom's *mass number* (*A*).

Chapter Summary, *Continued*

3. What are isotopes, and what is atomic weight?

- Atoms with identical numbers of protons and electrons but different numbers of neutrons are called *isotopes*.
- The atomic weight of an element is the weighted average mass of atoms of the element's naturally occurring isotopes measured in *atomic mass units* (amu).

Chapter Summary, *Continued*

4. How is the periodic table arranged?

- Elements are organized into the *periodic table*, consisting of 7 rows, or *periods*, and 18 columns, or *groups*.
- The two groups on the left side of the table and the six groups on the right are called the *main group elements*.
- The ten groups in the middle are the *transition metal groups*.
- The 14 groups pulled out and displayed below the main part of the table are the *inner transition metal groups*.
- Within a given group in the table, elements have the same number of valence electrons in their valence shell and similar electron configurations.

Chapter Summary, *Continued*

5. How are electrons arranged in atoms?

- The electrons surrounding an atom are grouped into *shells*. Within each shell, electrons are grouped into *subshells*, and within each subshell into *orbitals*.
- *s* orbitals are spherical; *p* orbitals are dumbbell-shaped.
- Each shell can hold a specific number of electrons. The first shell can hold 2 electrons, the second shell can hold 8 electrons, and the third shell can hold 18 electrons.
- The *electron configuration* of an element is predicted by assigning the element's electrons into orbitals.
- The electrons in the *valence shell* can be represented using electron-dot symbols.

Key Words

- Alkali metal
- Alkaline earth metal
- Atom
- Atomic mass unit (amu)
- Atomic number (*Z*)
- Atomic theory
- Atomic weight
- *d*-Block element
- Electron
- Electron configuration
- Electron-dot symbol
- *f*-Block element
- Group
- Halogen
- Inner transition metal element
- Isotopes

Key Words, *Continued*

- Main group element
- Mass number (*A*)
- Neutron
- Noble gas
- Nucleus
- Orbital
- *p*-Block element
- Period
- Proton
- *s*-Block element
- Shell (electron)
- Subatomic particles
- Subshell (electron)
- Transition metal element
- Valence electron
- Valence shell