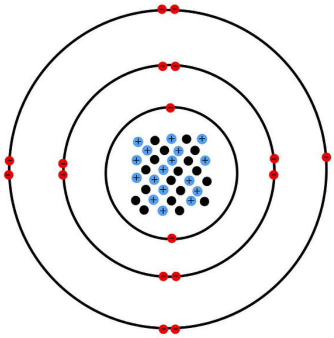
Reflect

Think about all the things in the world around you that radiate light energy—that is, things that glow or produce visible light. Some examples are light bulbs, fire, neon signs, lightning, and the stars in the sky (including the Sun). Though it may not always be obvious to us, the glow of a radiating object contains special information about its identity.

Why do some substances radiate light? What does this radiated light tell us about the substance?



The Bohr model of a chlorine atom shows the different energy levels at which an electron can orbit the atomic nucleus. The ring closest to the nucleus is n = 1, the middle ring is n = 2, and the outer orbital is n = 3. When the electron gains energy, it moves to a higher level; when it loses energy, it drops to a lower level.

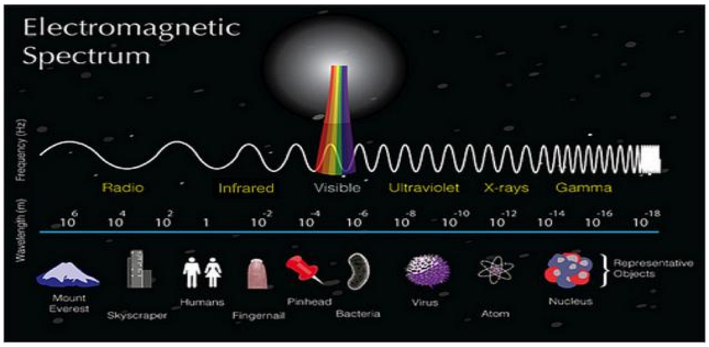
Electrons and Energy Levels in Atoms

Recall that the core of an atom is its nucleus. This is where positively charged protons are located, as well as neutrons, which have zero charge. Negatively charged electrons are located outside the atomic nucleus, where they occupy specific energy levels, or orbitals. The energy of an orbital is measured in electron volts, or eV. The Bohr model of the atom chlorine, shown at right, symbolically represents these orbitals. Electrons are typically located in the lowest available energy level, also known as the ground state. In the diagram on the right, the ground state corresponds to the level n = 1. However, sometimes an electron will gain energy and jump up to another energy level. This could happen if an electromagnetic wave strikes the electron. Alternatively, it could happen if the atom collides with another atom moving at high speed.

In order for an electromagnetic wave to increase the energy of an electron, the energy of the wave must be exactly equal to the energy difference between the electron's current energy level and another energy level. For example, consider a single electron in a hydrogen atom that exists at the ground state. The difference in energy between the levels n = 1 and n = 2 is 10.2 electron volts. If an electromagnetic wave with exactly 10.2 electron volts strikes the atom, the electron will absorb this energy and move to level n = 2. (If the amount of energy that strikes the atom is 10.1 electron volts or 10.3 electron volts, the electron will not move to level n = 2.) Alternatively, the difference in energy between levels n = 1 and n = 3 is 12.1 electron volts. So if an electromagnetic wave with exactly 12.1 electron volts strikes the atom, the electron will move to level n = 3. Note that the energy of an electromagnetic wave is directly proportional to its frequency and inversely proportional to its wavelength. So shorter waves have more energy than longer waves. These waves actually travel in "wave packets" or bundles called a photon. Each photon is composed of waves of just one wavelength.

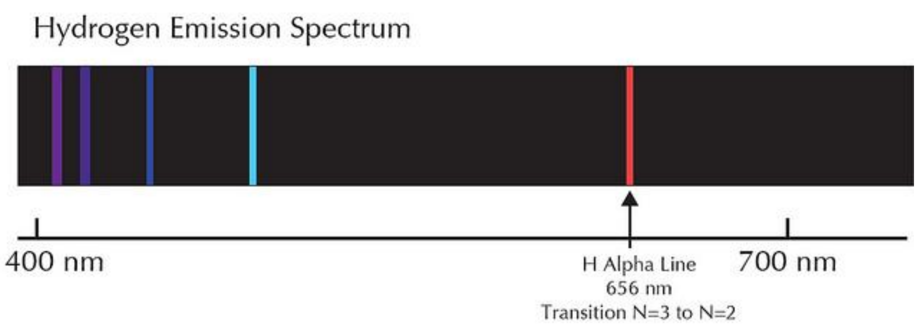
Electron Movement and Configuration

After an electron has moved to a higher energy level, it typically remains there for a very short time before spontaneously dropping back to a lower energy level, typically the ground state. As it drops, it releases the electromagnetic energy it had absorbed. So if the electron had absorbed energy of light corresponding to the red wavelength of visible light, it would emit red light when it returned to the ground state. If it drops to an intermediate energy level, it emits radiation with energy equal to the difference between the higher and the intermediate level. All radiation occurs when electrons drop from a higher energy level to a lower level and emit an electromagnetic wave.



The electromagnetic spectrum consists of all wavelengths of electromagnetic radiation. The wavelengths that humans can see make up the visible light part of the spectrum. As you can see, each color of visible light corresponds to a different wavelength. Red light waves have the longest wavelengths and lowest energies of all visible light.

The range of different waves that can be released when electrons fall from a higher energy level to a lower energy level in an atom is called the emission spectrum for that atom. Consider billions of hydrogen atoms in a region. A scientist could excite their electrons to different energy levels. In this case, there would be a specific set of light waves that these atoms would emit as the various electrons in different excited states eventually cascade down to their ground states. Scientists use spectroscopes to observe these waves. The emission spectrum for hydrogen is displayed below.



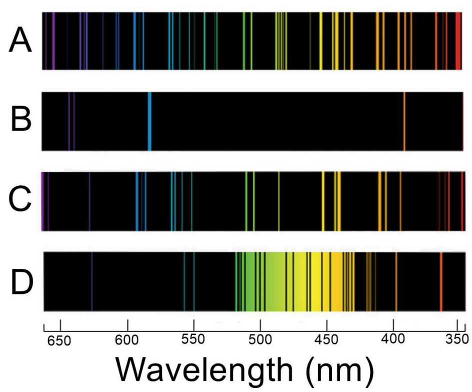
A spectroscope (or ... nto different wavelengths. So the light emitted from hot hydrogen gas would travel through a spectroscope, and the different wavelengths of light emitted from the atoms would bend (or diffract) at different angles onto the spectrum shown. Note that the visible bands of light shown on the spectrum are well defined, and their positions correspond to their specific wavelengths of light, measured in nanometers (nm).

Each element on the periodic table has a unique emission spectrum. This is because the particular ways in which electrons can fall to lower energy levels depends on the specific configuration of the atom. So a scientist can pass the light emitted by any glowing gas (or very hot object) through a spectroscope and observe its unique emission spectrum. In this way, emission spectra can be used to determine the identities of unknown elements on Earth or into space.

In order to identify the different atoms in a substance, the mass spectroscopist must ionize the substance. This can be done in many ways, but one way is by heating the substance to knock electrons out of orbit. This gives each atom a positive or negative charge. Charged atoms are ions. Then the mass spectroscopist separates the ions by charge, placing them in a chamber with an electromagnetic beam; the extent to which an ion deflects from the beam depends on its mass. In this way, the mass spectroscopist can measure and analyze masses and charges of unknown ions in a substance to determine the identity of the substances. One application of this mass spectroscopy is urine testing. If unusual values show up in the mass spectrum for a urine sample, the spectroscopist can compare this with the known mass spectra for different drugs to determine whether a drug is present in the urine.

Try Now

The emission spectra of five unknown elements are shown below. Study these spectra, then match each emission spectrum to the appropriate element in the table. (Note: Not every wavelength shown in the spectra is included in the table.)

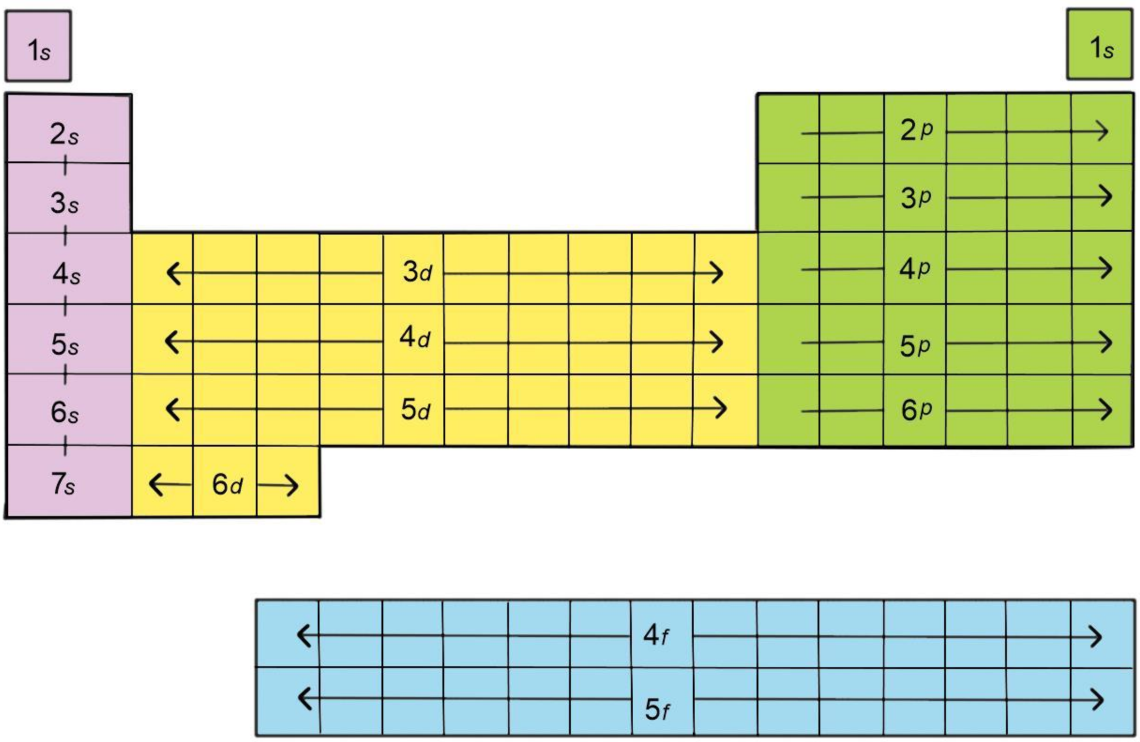


Light Emission

Element	Emission Wavelength (nm)	Letter
Neon (Ne)	660, 651, 640, 627, 616, 607, 588, 585, 540	
Mercury (Hg)	579, 576, 546, 492, 436	
Sodium (Na)	589	
Hydrogen (H)	656, 486, 434, 410, 397, 389, 384	
Helium (He)	668, 588, 492, 471, 447, 444, 439	

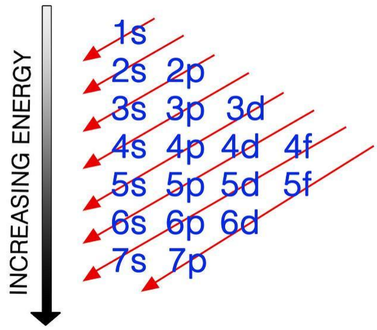
Electrons are located in energy levels.
The orbitals that hold electrons are in groups known as energy levels. Each group is a different distance from the nucleus. Electrons closer to the nucleus are lower in energy than electrons farther from the nucleus. The principle energy levels are denoted by the variable n, where n = 1, 2, 3, and so on. Energy levels are divided into four sublevels: the s, p, d, and f levels. These sublevels increase in energy: s < p < d < f. Each sublevel corresponds to a particular location on the periodic table. Likewise, each sublevel can contain a maximum number of electrons. A good way to remember the order or the sublevels is with the phrase: "**s**ome **p**igs **d**o **f**ly."

- **s**: The s sublevel corresponds to the first two groups (1A and 2A) on the periodic table. Only two electrons can occupy this sublevel. Each energy level has an s sublevel (1s, 2s, 3s, and so on).
- **p**: The p sublevel corresponds to groups 3A–8A on the periodic table. This sublevel can contain up to six electrons. Energy level n = 1 does not have a p sublevel (only 2p, 3p, and so on).
- **d**: The d sublevel corresponds to the d-block of the transition elements (groups 1B–10B). This sublevel can contain up to 10 electrons. Only the energy levels n = 3, 4, 5, and 6 have d sublevels (3d, 4d, 5d, and 6d).
- **f**: The f sublevel corresponds to the inner transition elements (lanthanide and actinide series). This sublevel can contain up to 14 electrons. Only the energy levels n = 4 and 5 can have f sublevels (4f and 5f).



Electron Movement and Configuration

These energy levels and sublevels can be used to write an electron configuration for each element. An electron configuration is a written way to show the location and number of electrons in an atom. To describe the location of an electron, you must name both the energy level (n = 1, 2, 3, ...) and the sublevel (s, p, d, or f) of the electron. The sublevels are filled in order from lowest to highest energy. This is referred to as the aufbau principle: The 1s orbital is filled first, followed by the 2s orbital, the 2p orbital, the 3s orbital, and so on. Let's look at an atom of chlorine (Cl), which has an atomic number of 17. Therefore, a neutral chlorine atom contains a total of 17 electrons. (Remember, in a neutral atom, the number of positive charges and negative charges are equal.) Because orbitals are filled sequentially from lower energy to higher energy, the electron configuration for chlorine is written as $1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^5$. The superscripts represent the number of electrons in that energy level. If you add the superscripts for this particular atom, you get 17 (2 + 2 + 6 + 2 + 5 = 17). This is a way to check to see if you completed the electron configuration correctly. Your atomic number or number of electrons in a neutral atom should add up all your superscripts in your electron configuration. For example, a ground state chlorine atom has an atomic number or 17; when you add the superscripts together, you will also have a sum of 17. Remember, each energy sublevel has a maximum number of electrons it can contain.



This diagram shows the order in which electrons fill orbitals. Starting at the top of the diagram (1s), follow each arrow down and to the left to determine when a particular orbital is filled.

What Do You Think?

Here are electron configurations for three different atoms. Can you identify each atom based on its electron configuration? Place the name of the element in the blank next to its electron configuration.

1. $1s^2$, $2s^2$, $2p^6$, $3s^1$ ______________________________
2. $1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^3$ ______________________________
3. $1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^6$, 4s2, $3d^5$ ______________________________
4. [Ne] $3s^2$, $3p^4$ ______________________________
5. K^+ (write the electron configuration of this potassium ion)

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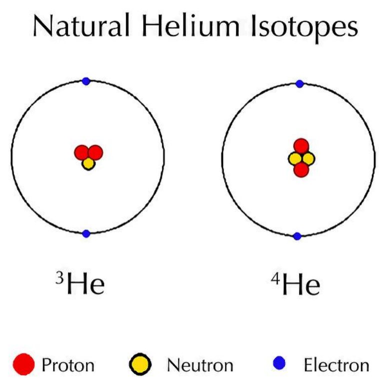
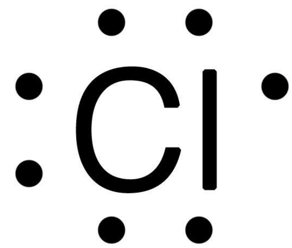
Look Out!

Isotopes are atoms of the same element that have different numbers of neutrons. There are two natural isotopes of helium (He). One isotope has one neutron, and the other isotope has two neutrons. Both isotopes have two protons. Therefore, the mass number—the total number of an atom's protons and neutrons—of the first helium isotope is three (two protons plus one neutron). The mass number of the second helium isotope is four (two protons plus two neutrons). Helium has only two isotopes, but most elements have more. Iron (Fe) has more than 25 isotopes, but many of these isotopes are not very stable. The most abundant iron isotope has 26 protons and 30 neutrons. Therefore, this isotope has a mass number of 56.



Lewis Valence Electron Dot Structure

Remember, only the valence electrons of an atom are responsible for interactions with other atoms. For this reason, scientists use the Lewis valence electron dot structure as a model to represent the valence electrons in an atom. In this model, the chemical symbol of an element is surrounded by dots. Each dot represents a valence electron. In general, atoms with eight electrons in their valence shell are the most stable. (This is called the octet rule.) Exceptions include hydrogen and helium; atoms of these elements can have at most two valence electrons. In a Lewis dot structure, the electrons, or dots, surround the chemical symbol in pairs. Each "side" of the chemical symbol (above, below, left, right) can contain an electron pair, for a total of eight valence electrons. Each of the first four valence electrons occupies an unpaired space. Below you can see an example of Lewis dot structure.



The Lewis dot structure for chlorine (Cl) shows three electron pairs and one unpaired electron, for a total of seven electrons.

Atomic Structure

Because each element in a group of the periodic table has the same number of valence electrons, the Lewis dot structure of all elements in a group is the same. For example, both fluorine (F) and chlorine (Cl) are found in Group 7A on the periodic table, so their Lewis dot structures contain seven electrons, as you see at right. Also, elements with the same number of valence electrons will tend to act or react the same. They will also tend to have similar chemical properties. The electrons shown in Lewis dot structures represent electrons in the s and p orbitals, so Lewis dot structures are not often used to model transition elements.

Try Now

1. Electrons occupy the lowest energy orbital first, then move to the next one and so on (the aufbau principle).
2. Orbitals are considered to be in the same shell if they have the same first number (no matter the order of the filling).
3. An atom will gain or lose electrons in order to have eight electrons in its outer shell (the octet rule).
4. The outer shell is the highest numbered shell that has electrons in it. Only s and p orbitals are part of the outer shell.

An atom has the tendency to lose electrons (to another atom) or to gain electrons (from another atom) in order to make the outer shell complete with eight electrons. Atoms with a complete outer shell (eight electrons) are considered stable. Some atoms naturally have eight electrons in their outer shell and are very stable. (Helium is the exception, being stable with two electrons in its outer shell.) Complete the following chart:

Element	Atomic Number	Electron Configuration	Lewis Dot Structure
Helium (He)			
Carbon (C)	6		
		$1s^{2,}$ $2s^2$,$2p^6$, $3s^2$, $3p^6$	
Calcium (Ca)			$\overset{..}{Ca}$
		[Ar] $4s^2$, $3d^6$	

Connecting With Your Child

To help your child learn more about emission spectra, have him or her construct a basic spectrometer to observe the emission spectra of light in everyday life.

To do this, you will need the following materials:

- Empty cereal box (taped shut)
- Sharp knife or razor
- Compact disc
- Masking tape
- Protractor

The compact disc has diffraction grating on the shiny bottom surface. When light strikes this surface, it spreads out into the different wavelengths emitted from the light source. This allows an observer to see the emission spectra from a particular light source.

How to Construct the Spectrometer

- Have your child cut a five centimeter by five centimeter square into the top of the box, on the right corner. This will be the hole through which you will observe the emission spectra.
- Then assist in cutting an eight centimeter diagonal slit from the right corner toward the center of the box along the front face of the box. The slit should make a 30 degree angle with the right side of the box. An identical slit should be cut from the right corner toward the center of the box along the back face of the box.
- Slide the compact disc into the slits so that the shiny surface faces upward. You should be able to see this shiny surface when you look down through the hole cut into the top of the box.
- Next, have your child cut a five centimeter by five centimeter hole into the top-left side of the box, opposite the compact disc that was inserted on the right side of the box. When you look through this hole, you should be able to see the compact disc at the other end of the box.
- Then use masking tape to cover this hole so there is a tiny slit, one millimeter in width, through which light can enter the cereal box.

How Does It Work?

The spectrometer works when light from a source passes through this tiny slit and then diffracts along the compact disc. Your child can then observe the emission spectra by looking through the hole in the top of the box and observing the pattern along the compact disc. Have your child hold the slit-end of the box to different light sources, such as an incandescent light, fluorescent light, plasma screen television, computer screen, neon signs, and so forth.

Here are some questions to discuss with your child:

1. Why does light spread out into different colors along the compact disc surface?
2. What do these colors reveal about the light source?
3. How does the emission spectrum differ from one light source to another?
4. Did you observe any light sources that show only one wavelength of light through the spectrometer? If not, why do you think this is? If so, how is this source different from others?