Chapter 9
Electrons in Atoms and the Periodic Table

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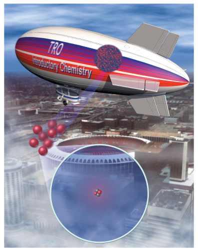
Chem 118
Introductory Chemistry



Map: Introductory Chemistry (Tro) https://chem.libretexts.org/@go/page/45050 (accessed Mar 25, 2022).

Blimps

- Blimps float because they are filled with a gas that is less dense than the surrounding air.
- Early blimps used the gas **hydrogen**, however, hydrogen's flammability lead to the Hindenburg disaster.
- Blimps now use helium gas, which is not flammable. In fact, it doesn't undergo any chemical reactions.
- This chapter investigates models of the atom we use to explain the differences in the properties of the elements.

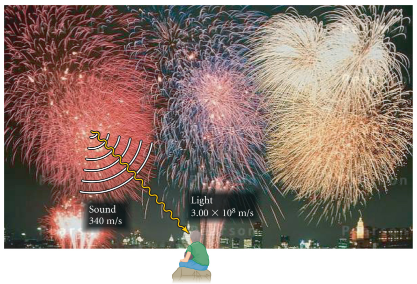
Classical View of the Universe

- Since the time of the ancient Greeks, the stuff of the physical universe has been classified as either matter or energy.
- We define matter as the stuff of the universe that has mass and volume.
  - ✓ Therefore, energy is the stuff of the universe that doesn't have mass and volume.
- We know from our examination of matter that it is ultimately composed of particles, and its the properties of those particles that determine the properties we observe.
- Energy, therefore, should not be composed of particles. In fact, the thing that all energy has in common is that it travels in **waves**.

The Nature of Light—Its Wave Nature

- Light is one of the forms of energy.
- Light is a form of **electromagnetic radiation**.
- Electromagnetic radiation is made of waves.
- Electromagnetic radiation moves through space like waves move across the surface of a pond

Speed of Energy Transmission



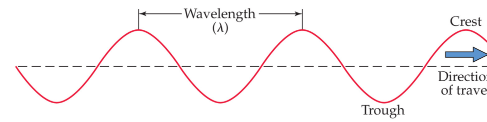
Electromagnetic Waves

- Every wave has four characteristics that determine its properties:
  - ✓ wave speed,
  - ✓ height (amplitude),
  - ✓ length,
  - ✓ number of wave peaks that pass in a given time.
- All electromagnetic waves move through space at the same, constant speed.
  - ✓ 3.00×10^8 meters per second in a vacuum = **The speed of light, c**.

Characterizing Waves

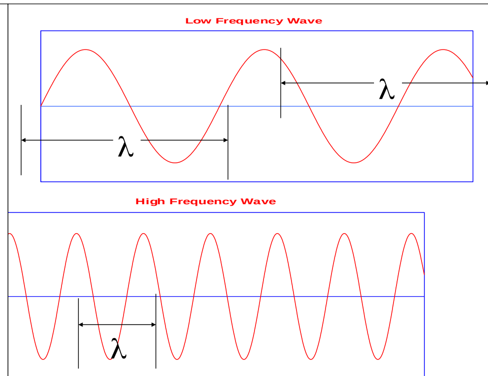
- The **amplitude** is the height of the wave.
  - ✓ The distance from node to crest.
    - ➢ Or node to trough.
  - ✓ The amplitude is a measure of how intense the light is—the larger the amplitude, the brighter the light.
- The **wavelength (λ)** is a measure of the distance covered by the wave.
  - ✓ The distance from one crest to the next.
    - ➢ Or the distance from one trough to the next, or the distance between alternate nodes.
  - ✓ Usually measured in nanometers.
    - ➢ $1 \text{ nm} = 1 \times 10^{-9}$ m

Electromagnetic Waves



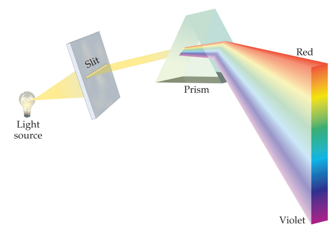
Characterizing Waves

- The **frequency (ν)** is the number of waves that pass a point in a given period of time.
  - ✓ The number of waves = number of cycles.
  - ✓ Units are hertz (Hz), or cycles/s = s^{-1}.
    - ➢ $1 \text{ Hz} = 1 \text{ s}^{-1}$
- The total energy is proportional to the amplitude and frequency of the waves.
  - ✓ The larger the wave amplitude, the more force it has.
  - ✓ The more frequently the waves strike, the more total force there is.



The Electromagnetic Spectrum

- Light passed through a prism is separated into all its colors. This is called a **continuous spectrum**.
- The color of the light is determined by its wavelength.



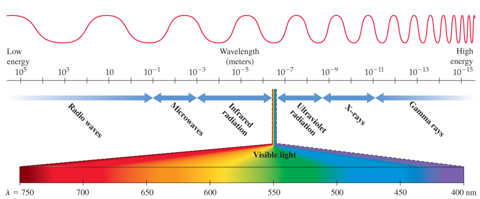
Color

- The color of light is determined by its wavelength.
  - ✓ Or frequency.
- White light is a mixture of all the colors of visible light.
  - ✓ A spectrum.
  - ✓ **R**ed**O**range**Y**ellow**G**reen**B**lue**I**ndigo**V**iolet.
- When an object absorbs some of the wavelengths of white light while reflecting others, it appears colored.
  - ✓ The observed color is predominantly the colors reflected.

Types of Electromagnetic Radiation

- Classified by the Wavelength
  - ✓ Radiowaves = $\lambda > 0.01$ m.
    - ➢ Low frequency and energy.
  - ✓ Microwaves = $10^{-4}\text{ m} < \lambda < 10^{-2}\text{ m}$.
  - ✓ Infrared (IR) = $8 \times 10^{-7}\text{ m} < \lambda < 10^{-5}\text{ m}$.
  - ✓ Visible = $4 \times 10^{-7}\text{ m} < \lambda < 8 \times 10^{-7}\text{ m}$.
    - ➢ ROYGBIV.
  - ✓ Ultraviolet (UV) = $10^{-8}\text{ m} < \lambda < 4 \times 10^{-7}\text{ m}$.
  - ✓ X-rays = $10^{-10}\text{ m} < \lambda < 10^{-8}\text{ m}$.
  - ✓ Gamma rays = $\lambda < 10^{-10}$ m.
    - ➢ High frequency and energy.

Electromagnetic Spectrum



Particles of Light

- Scientists in the early 20th century showed that electromagnetic radiation was composed of particles we call **photons**.
  - ✓ Max Planck and Albert Einstein.
  - ✓ Photons are particles of light energy.
- Each wavelength of light has photons that have a different amount of energy.

The Electromagnetic Spectrum and Photon Energy

- Short wavelength light has photons with high energy.
- High frequency light has photons with high energy.
  - ✓ Radiowave photons have the lowest energy.
  - ✓ Gamma ray photons have the highest energy.
- High-energy electromagnetic radiation can potentially damage biological molecules.
  - ✓ Ionizing radiation.

Order the Following Types of Electromagnetic Radiation: Microwaves, Gamma Rays, Green Light, Red Light, Ultraviolet Light

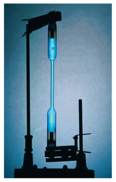
- By wavelength (short to long).
- By frequency (low to high).
- By energy (least to most).

Order the Following Types of Electromagnetic Radiation: Microwaves, Gamma Rays, Green Light, Red Light, Ultraviolet Light, Continued

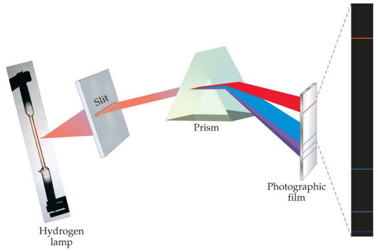
- By wavelength (short to long).
  Gamma < UV < green < red < microwaves.
- By frequency (low to high).
  Microwaves < red < green < UV < gamma.
- By energy (least to most).
  Microwaves < red < green < UV < gamma.

Light's Relationship to Matter

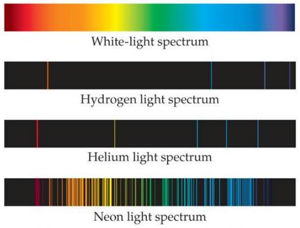
- Atoms can acquire extra energy, but they must eventually release it.
- When atoms emit energy, it usually is released in the form of light.
- However, atoms don't emit all colors, only very specific wavelengths.
  - ✓ In fact, the spectrum of wavelengths can be used to identify the element.

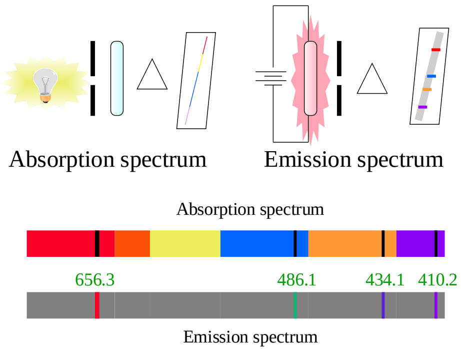


Emission Spectrum



Spectra



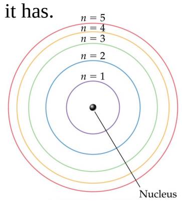


The Bohr Model of the Atom

- The nuclear model of the atom does not explain how the atom can gain or lose energy.
- Neils Bohr developed a model of the atom to explain how the structure of the atom changes when it undergoes energy transitions.
- Bohr's major idea was that the energy of the atom was **quantized**, and that the amount of energy in the atom was related to the electron's position in the atom.
  - ✓ Quantized means that the atom could only have very specific amounts of energy.

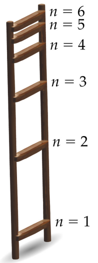
The Bohr Model of the Atom: Electron Orbits

- In the Bohr model, electrons travel in orbits around the nucleus.
  - ✓ More like shells than planet orbits.
- The farther the electron is from the nucleus the more energy it has.



The Bohr Model of the Atom: Orbits and Energy, Continued

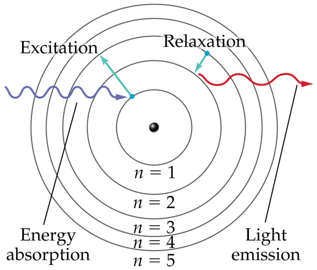
- Each orbit has a specific amount of energy.
- The energy of each orbit is characterized by an integer—the larger the integer, the more energy an electron in that orbit has and the farther it is from the nucleus.
  - ✓ The integer, **n**, is called a **quantum number**.



The Bohr Model of the Atom: Energy Transitions

- When the atom gains energy, the electron leaps from a lower energy orbit to one that is further from the nucleus.
  - ✓ However, during that "quantum leap" it doesn't travel through the space between the orbits, it just disappears from the lower orbit and appears in the higher orbit.
- When the electron leaps from a higher energy orbit to one that is closer to the nucleus, energy is emitted from the atom as a photon of light—a **quantum** of energy.

The Bohr Model of the Atom



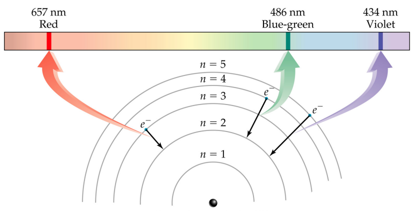
The Bohr Model of the Atom: Ground and Excited States

- In the Bohr model of hydrogen, the lowest amount of energy hydrogen's one electron can have corresponds to being in the $n = 1$ orbit. We call this its **ground state.**
- When the atom gains energy, the electron leaps to a higher energy orbit. We call this an **excited state**.
- The atom is less stable in an excited state and so it will release the extra energy to return to the ground state.
  - ✓ Either all at once or in several steps.

The Bohr Model of the Atom: Hydrogen Spectrum

- Every hydrogen atom has identical orbits, so every hydrogen atom can undergo the same energy transitions.
- However, since the distances between the orbits in an atom are not all the same, no two leaps in an atom will have the same energy.
  - ✓ The closer the orbits are in energy, the lower the energy of the photon emitted.
  - ✓ Lower energy photon = longer wavelength.
- Therefore, we get an emission spectrum that has a lot of lines that are unique to hydrogen.

The Bohr Model of the Atom: Hydrogen Spectrum, Continued

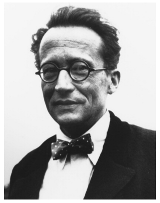


The Bohr Model of the Atom: Success and Failure

- The mathematics of the Bohr model very accurately predicts the spectrum of hydrogen.
- However, its mathematics fails when applied to multi-electron atoms.
  - ✓ It cannot account for electron-electron interactions.
- A better theory was needed.

The Quantum-Mechanical Model of the Atom

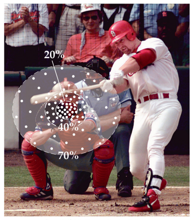
- Erwin Schrödinger applied the mathematics of probability and the ideas of quantizing energy to the physics equations that describe waves, resulting in an equation that predicts the **probability** of finding an electron with a particular amount of energy at a particular location in the atom.



The Quantum-Mechanical Model: Orbitals

- The result is a map of regions in the atom that have a particular probability for finding the electron.
- An **orbital** is a region where we have a very high probability of finding the electron when it has a particular amount of energy.
  - ✓Generally set at 90 or 95%.

Orbits vs. Orbitals Pathways vs. Probability

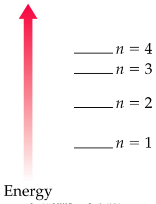


Wave–Particle Duality

- We've seen that light has the characteristics of waves and particles (photons) at the same time—how we view it depends on the application.
- In the same way, electrons have the characteristics of both particles and waves at the same time.
- This makes it impossible to predict the path of an electron in an atom.

The Quantum-Mechanical Model: Quantum Numbers

- In Schrödinger's wave equation, there are 3 integers, called **quantum numbers**, that quantize the energy.
- The **principal quantum number, *n***, specifies the main energy level for the orbital.



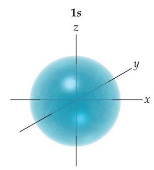
The Quantum-Mechanical Model: Quantum Numbers, Continued

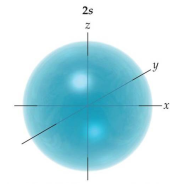
- Each principal energy shell has one or more subshells.
  - ✓The number of subshells = the principal quantum number.
- The quantum number that designates the subshell is often given a letter.
  - ✓*s, p, d, f*.
- Each kind of sublevel has orbitals with a particular shape.
  - ✓The shape represents the probability map.
    - ➢90% probability of finding electron in that region.

Shells and Subshells

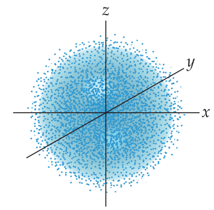
Shell	# of subshells	Letters specifying subshells
$n = 4$	4	*s* *p* *d* *f*
$n = 3$	3	*s* *p* *d*
$n = 2$	2	*s* *p*
$n = 1$	1	*s*

How Does the 1*s* Subshell Differ from the 2*s* Subshell?

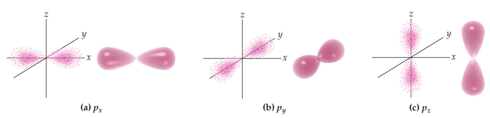




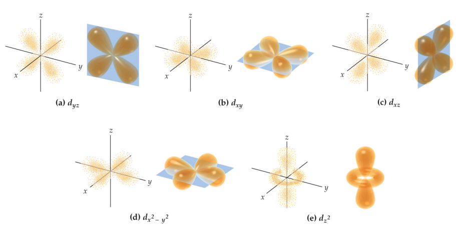
Probability Maps and Orbital Shape: *s* Orbitals



Probability Maps and Orbital Shape: *p* Orbitals



Probability Maps and Orbital Shape: *d* Orbitals



Subshells and Orbitals

- The subshells of a principal shell have slightly different energies.
  - ✓The subshells in a shell of H all have the same energy, but for multielectron atoms the subshells have different energies.
  - ✓$s < p < d < f$.
- Each subshell contains one or more orbitals.
  - ✓*s* subshells have 1 orbital.
  - ✓*p* subshells have 3 orbitals.
  - ✓*d* subshells have 5 orbitals.
  - ✓*f* subshells have 7 orbitals.

The Quantum-Mechanical Model: Energy Transitions

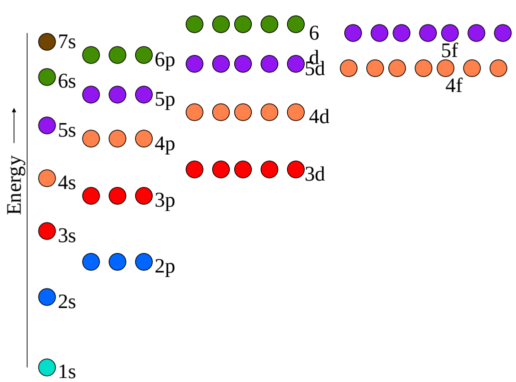
- As in the Bohr model, atoms gain or lose energy as the electron leaps between orbitals in different energy shells and subshells.
- The **ground state** of the electron is the lowest energy orbital it can occupy.
- Higher energy orbitals are **excited states**.

The Bohr Model vs. the Quantum-Mechanical Model

- Both the Bohr and quantum-mechanical models predict the spectrum of hydrogen very accurately.
- Only the quantum-mechanical model predicts the spectra of multi-electron atoms.

Electron Configurations

- The distribution of electrons into the various energy shells and subshells in an atom in its ground state is called its **electron configuration**.
- Each energy shell and subshell has a maximum number of electrons it can hold.
  - ✓$s = 2$, $p = 6$, $d = 10$, $f = 14$.
  - ✓Based on the number of orbitals in the subshell.
- We place electrons in the energy shells and subshells in order of energy, from low energy up.
  - ✓Aufbau principle.



Filling an Orbital with Electrons

- Each orbital may have a maximum of 2 electrons.
  - ✓Pauli Exclusion principle.
- Electrons spin on an axis.
  - ✓Generating their own magnetic field.
- When two electrons are in the same orbital, they must have opposite spins.
  - ✓So their magnetic fields will cancel.

Orbital Diagrams

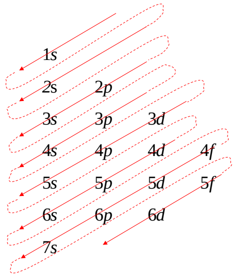
- We often represent an orbital as a square and the electrons in that orbital as arrows.
  - ✓ The direction of the arrow represents the spin of the electron.

Unoccupied orbital	Orbital with 1 electron	Orbital with 2 electrons
□	↑	↑↓

Order of Subshell Filling in Ground State Electron Configurations

Start by drawing a diagram putting each energy shell on a row and listing the subshells (*s, p, d, f*) for that shell in order of energy (left to right).

Next, draw arrows through the diagonals, looping back to the next diagonal each time.



Filling the Orbitals in a Subshell with Electrons

- Energy shells fill from lowest energy to highest.
- Subshells fill from lowest energy to highest.
  - ✓ *s* → *p* → *d* → *f*
- Orbitals that are in the same subshell have the same energy.
- When filling orbitals that have the same energy, place one electron in each before completing pairs.
  - ✓ Hund's rule.

Electron Configuration of Atoms in their Ground State

- The electron configuration is a listing of the subshells in order of filling with the number of electrons in that subshell written as a superscript.

Kr = 36 electrons = $1s^22s^22p^63s^23p^64s^23d^{10}4p^6$

- A short-hand way of writing an electron configuration is to use the symbol of the previous noble gas in [] to represent all the inner electrons, then just write the last set.

Rb = 37 electrons = $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^1$ = [Kr]$5s^1$

Electron Configurations

Symbol	Number of electrons	Electron configuration	Orbital diagram
Li	3	$1s^22s^1$	1s: ↑↓ 2s: ↑
Be	4	$1s^22s^2$	1s: ↑↓ 2s: ↑↓
B	5	$1s^22s^22p^1$	1s: ↑↓ 2s: ↑↓ 2p: ↑ □ □
C	6	$1s^22s^22p^2$	1s: ↑↓ 2s: ↑↓ 2p: ↑ ↑ □

Example—Write the Ground State Orbital Diagram and Electron Configuration of Magnesium.

1. Determine the atomic number of the element from the periodic table.
   - ✓ This gives the number of protons and electrons in the atom.

Mg Z = 12, so Mg has 12 protons and 12 electrons.

Example—Write the Ground State Orbital Diagram and Electron Configuration of Magnesium, Continued.

2. Draw 9 boxes to represent the first 3 energy levels ***s*** and ***p*** orbitals.

1s	2s	2p	3s	3p
□	□	□ □ □	□	□ □ □

Example—Write the Ground State Orbital Diagram and Electron Configuration of Magnesium, Continued.

3. Add one electron to each box in a set, then pair the electrons before going to the next set until you use all the electrons.
   - When paired, put in opposite arrows.

1s	2s	2p	3s	3p
↑↓	↑↓	↑↓ ↑↓ ↑↓	↑↓	□ □ □

Example—Write the Ground State Orbital Diagram and Electron Configuration of Magnesium, Continued.

4. Use the diagram to write the electron configuration.
   - ✓ Write the number of electrons in each set as a superscript next to the name of the orbital set.

$1s^22s^22p^63s^2$ = [Ne]$3s^2$

1s	2s	2p	3s	3p
↑↓	↑↓	↑↓ ↑↓ ↑↓	↑↓	□ □ □

Example—Write the Full Ground State Orbital Diagram and Electron Configuration of Manganese

Mn Z = 25, therefore 25 e⁻

s subshell holds 2 e⁻

p subshell holds 6 e⁻

d subshell holds 10 e⁻

f subshell holds 14 e⁻

Based on the order of subshell filling, we will need the first 7 subshells

2 e⁻ · +2 = 4e⁻ · +6 +2 = 12e⁻ · +6 +2 = 20e⁻ · +10 = 30e⁻

Therefore the electron configuration is $1s^22s^22p^63s^23p^64s^23d^5$

Practice—Write the Full Ground State Orbital Diagram and Electron Configuration of Potassium

K Z = 19, therefore 19 e⁻

s subshell holds 2 e⁻

p subshell holds 6 e⁻

d subshell holds 10 e⁻

f subshell holds 14 e⁻

2 e⁻ · +2 = 4e⁻ · +6 +2 = 12e⁻ · +6 +2 = 20e⁻

Therefore the electron configuration is $1s^22s^22p^63s^23p^64s^1$

Based on the order of subshell filling, we will need the first 6 subshells

Example—Write the Full Ground State Orbital Diagram and Electron Configuration of Sc^{3+}

Sc Z = 21, therefore 21 e⁻

therefore Sc^{3+} has 18 e⁻

s subshell holds 2 e⁻

p subshell holds 6 e⁻

d subshell holds 10 e⁻

f subshell holds 14 e⁻

2 e⁻ · +2 = 4e⁻ · +6 +2 = 12e⁻ · +6 = 18e⁻

Therefore the electron configuration is $1s^22s^22p^63s^23p^6$

Based on the order of subshell filling, we will need the first 5 subshells

Practice—Write the Full Ground State Orbital Diagram and Electron Configuration of F^-

F Z = 9, therefore 9 e⁻

therefore F^- has 10 e⁻

s subshell holds 2 e⁻

p subshell holds 6 e⁻

d subshell holds 10 e⁻

f subshell holds 14 e⁻

2 e⁻ · +2 = 4e⁻ · +6 +2 = 12e⁻

Therefore the electron configuration is $1s^22s^22p^6$

Based on the order of subshell filling, we will need the first 3 subshells.

Valence Electrons

- The electrons in all the subshells with the highest principal energy shells are called the **valence electrons**.
- Electrons in lower energy shells are called **core electrons**.
- Chemists have observed that one of the most important factors in the way an atom behaves, both chemically and physically, is the number of valence electrons.

Valence Electrons, Continued

Rb = 37 electrons = $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^1$

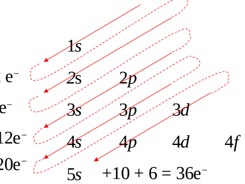
- The highest principal energy shell of Rb that contains electrons is the 5th, therefore, Rb has 1 valence electron and 36 core electrons.

Kr = 36 electrons = $1s^22s^22p^63s^23p^64s^23d^{10}4p^6$

- The highest principal energy shell of Kr that contains electrons is the 4th, therefore, Kr has 8 valence electrons and 28 core electrons.

Practice—Determine the Number and Types of Valence Electrons in an As Atom

As Z = 33, therefore 33 e⁻.



The highest occupied principal energy level is the 4th.

The valence electrons are 4*s* and 4*p* and there are 5 total.

Therefore, the electron configuration is $1s^22s^22p^63s^23p^64s^23d^{10}4p^3$.

Electron Configurations and the Periodic Table

1A							8A
1 **H** $1s^1$	2A	3A	4A	5A	6A	7A	2 **He** $1s^2$
3 **Li** $2s^1$	4 **Be** $2s^2$	5 **B** $2s^22p^1$	6 **C** $2s^22p^2$	7 **N** $2s^22p^3$	8 **O** $2s^22p^4$	9 **F** $2s^22p^5$	10 **Ne** $2s^22p^6$
11 **Na** $3s^1$	12 **Mg** $3s^2$	13 **Al** $3s^23p^1$	14 **Si** $3s^23p^2$	15 **P** $3s^23p^3$	16 **S** $3s^23p^4$	17 **Cl** $3s^23p^5$	18 **Ar** $3s^23p^6$

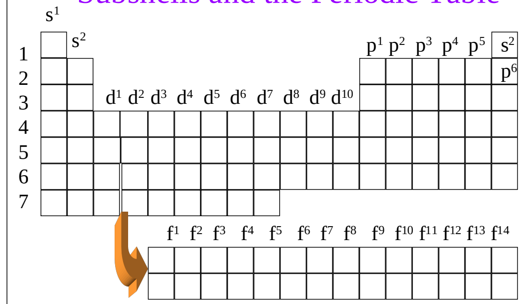
Electron Configurations from the Periodic Table

- Elements in the same period (row) have valence electrons in the same principal energy shell.
- The number of valence electrons increases by one as you progress across the period.
- Elements in the same group (column) have the same number of valence electrons and the valence electrons are in the same type of subshell.

Electron Configuration and the Periodic Table

- Elements in the same column have similar chemical and physical properties because their valence shell electron configuration is the same.
- The number of valence electrons for the main group elements is the same as the group number.

Subshells and the Periodic Table



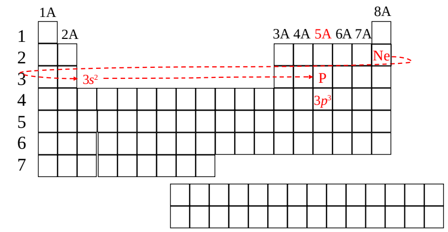
Electron Configuration from the Periodic Table

- The inner electron configuration is the same as the noble gas of the preceding period.
- To get the outer electron configuration from the preceding noble gas, loop through the next period, marking the subshells as you go, until you reach the element.
  - ✓ The valence energy shell = the period number.
  - ✓ The *d* block is always one energy shell below the period number and the *f* is two energy shells below.

Periodic Table and Valence Electrons

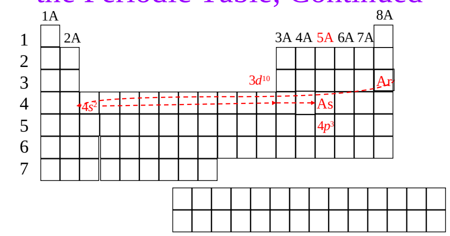
- For the main group elements, the number of valence electrons is the same as the column number.
  - ✓ Except for He.
- For the transition elements, the number of valence electrons is usually 2.
  - ✓ There are some elements whose electron configurations do not exactly fit our pattern.
  - ✓ Because as we traverse the transition metals we are putting electrons into a lower principal energy shell.

Electron Configuration from the Periodic Table



P = $[Ne]3s^23p^3$
P has 5 valence electrons.

Electron Configuration from the Periodic Table, Continued



As = $[Ar]4s^23d^{10}4p^3$
As has 5 valence electrons.

Practice—Use the Periodic Table to Write the Short Electron Configuration and Orbital Diagram for Each of the Following and Determine the Number of Valence Electrons.

- Na (at. no. 11).
- Te (at. no. 52).

Practice—Use the Periodic Table to Write the Short Electron Configuration and Orbital Diagram for Each of the Following and Determine the Number of Valence Electrons, Continued.

- Na (at. no. 11). $[Ne]3s^1$ 1 valence electron

  [↑]
  3s

- Te (at. no. 52). $[Kr]5s^24d^{10}5p^4$ 6 valence electrons

  [↑↓] [↑↓][↑↓][↑↓][↑↓][↑↓] [↑↓][↑][↑]
  5s 4d 5p

The Explanatory Power of the Quantum-Mechanical Model

- The properties of the elements are largely determined by the number of valence electrons they contain.
- Since elements in the same column have the same number of valence electrons, they show similar properties.
- Since the number of valence electrons increases across the period, the properties vary in a regular fashion.

The Noble Gas Electron Configuration

- The noble gases have 8 valence electrons.
  - ✓ Except for He, which has only 2 electrons.
- We know the noble gases are especially non-reactive.
  - ✓ He and Ne are practically inert.
- The reason the noble gases are so non-reactive is that the electron configuration of the noble gases is especially stable.

Noble gases 18 8A
2 **He** $1s^2$
10 **Ne** $2s^22p^6$
18 **Ar** $3s^23p^6$
36 **Kr** $4s^24p^6$
54 **Xe** $5s^25p^6$
86 **Rn** $6s^26p^6$

Everyone Wants to Be Like a Noble Gas! The Alkali Metals

- The alkali metals have one more electron than the previous noble gas.
- In their reactions, the alkali metals tend to lose their extra electron, resulting in the same electron configuration as a noble gas.
  - ✓ Forming a cation with a 1+ charge.

Alkali metals 1 1A
3 **Li** $2s^1$
11 **Na** $3s^1$
19 **K** $4s^1$
37 **Rb** $5s^1$
55 **Cs** $6s^1$
87 **Fr** $7s^1$

Everyone Wants to Be Like a Noble Gas! The Halogens

- The electron configurations of the halogens all have one fewer electron than the next noble gas.
- In their reactions with metals, the halogens tend to gain an electron and attain the electron configuration of the next noble gas.
  - ✓ Forming an anion with charge 1−.
- In their reactions with nonmetals, they tend to share electrons with the other nonmetal so that each attains the electron configuration of a noble gas.

Halogens 17 7A
9 **F** $2s^22p^5$
17 **Cl** $3s^23p^5$
35 **Br** $4s^24p^5$
53 **I** $5s^25p^5$
85 **At** $6s^26p^5$

78

Everyone Wants to Be Like a Noble Gas!

- As a group, the alkali metals are the most reactive metals.
  - ✓ They react with many things and do so rapidly.
- The halogens are the most reactive group of nonmetals.
- One reason for their high reactivity is the fact that they are only one electron away from having a very stable electron configuration.
  - ✓ The same as a noble gas.

Stable Electron Configuration and Ion Charge

- Metals form cations by losing valence electrons to get the same electron configuration as the previous noble gas.
- Nonmetals form anions by gaining valence electrons to get the same electron configuration as the next noble gas.

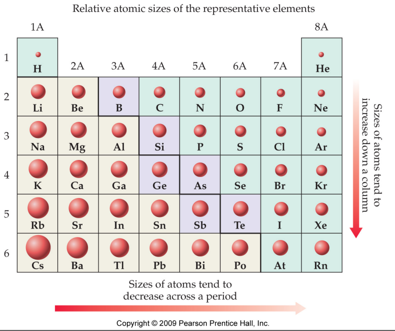
Atom	Atom's electron config	Ion	Ion's electron config
Na	$[Ne]3s^1$	Na^+	[Ne]
Mg	$[Ne]3s^2$	Mg^{2+}	[Ne]
Al	$[Ne]3s^23p^1$	Al^{3+}	[Ne]
O	$[He]2s2p^4$	O^{2-}	[Ne]
F	$[He]2s^22p^5$	F^-	[Ne]

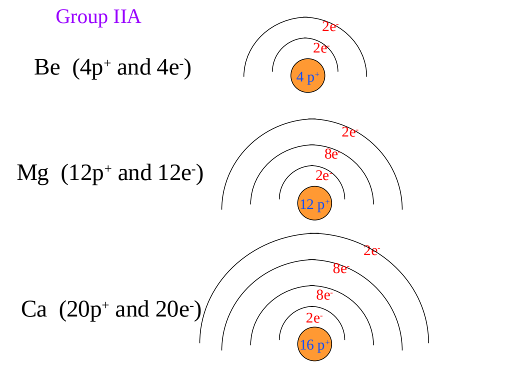
Periodic Trends in the Properties of the Elements

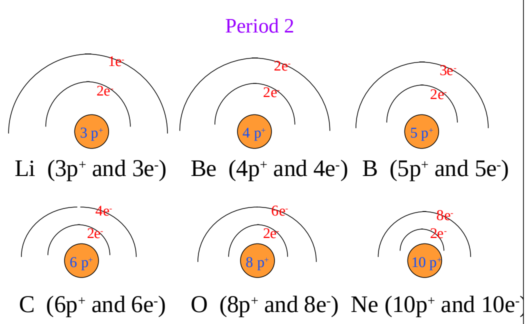
Trends in Atomic Size

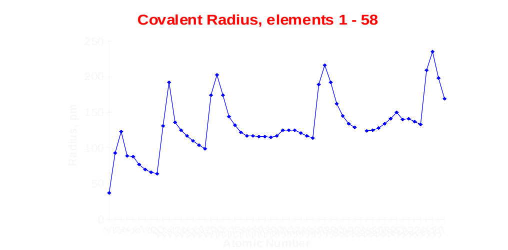
- Either volume or radius.
  - ✓Treat atom as a hard marble.
- As you traverse down a column on the periodic table, the size of the atom **increases**.
  - ✓Valence shell farther from nucleus.
  - ✓Effective nuclear charge fairly close.
- As you traverse left to right across a period, the size of the atom **decreases**.
  - ✓Adding electrons to same valence shell.
  - ✓Effective nuclear charge increases.
  - ✓Valence shell held closer.

Trends in Atomic Size, Continued



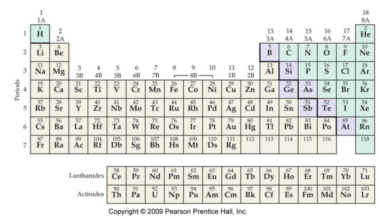






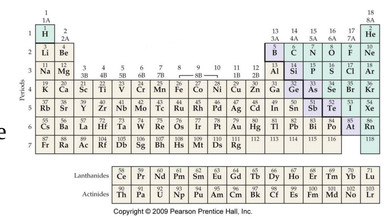
Example 9.6 – Choose the Larger Atom in Each Pair

- C or O
- Li or K
- C or Al
- Se or I?



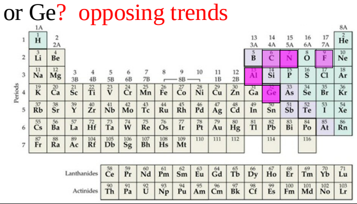
Practice—Choose the Larger Atom in Each Pair.

1. N or F
2. C or Ge
3. N or Al
4. Al or Ge

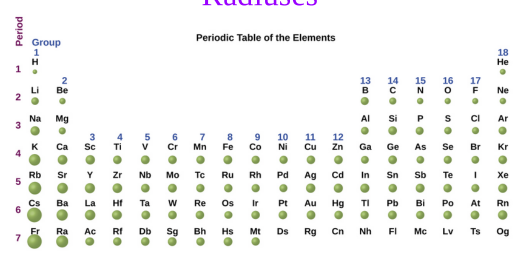


Practice—Choose the Larger Atom in Each Pair, Continued.

1. N or F, N is further left
2. C or Ge, Ge is further down
3. N or Al, Al is further down & left
4. Al or Ge? opposing trends



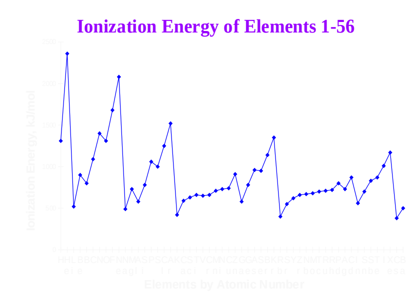
Periodic Table of Atomic Radiuses

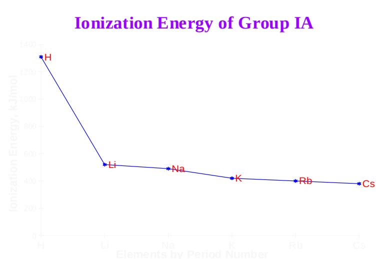


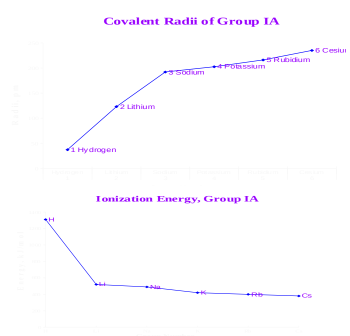
https://openstax.org/books/chemistry-2e/pages/6-5-periodic-variations-in-element-properties

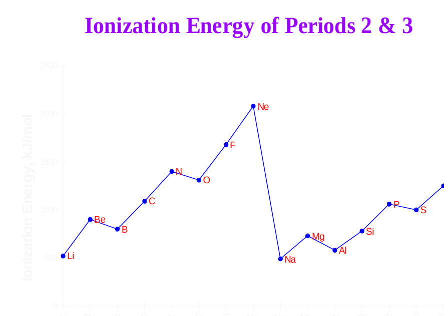
Ionization Energy

- Minimum energy needed to remove an electron from an atom.
  - ✓Gas state.
  - ✓Endothermic process.
  - ✓Valence electron easiest to remove.
  - ✓$M(g)$ + 1st IE → $M^{1+}(g)$ + 1 e-
  - ✓$M^{+1}(g)$ + 2nd IE → $M^{2+}(g)$ + 1 e-
    - ➢First ionization energy = energy to remove electron from neutral atom; 2nd IE = energy to remove from +1 ion; etc.





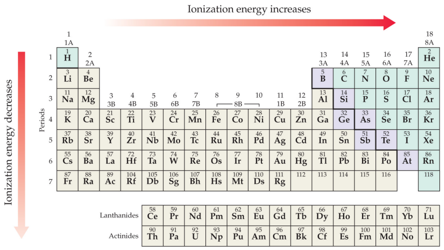




Trends in Ionization Energy

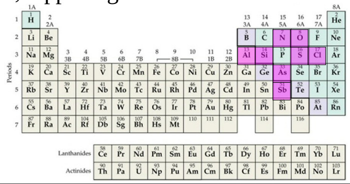
- As atomic radius increases, the ionization energy (IE) generally decreases.
  - ✓Because the electron is closer to the nucleus.
- 1st IE < 2nd IE < 3rd IE …
- As you traverse down a column, the IE gets **smaller**.
  - ✓Valence electron farther from nucleus.
- As you traverse left to right across a period, the IE gets **larger**.
  - ✓Effective nuclear charge increases.

Trends in Ionization Energy, Continued



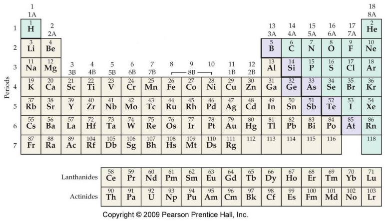
Example—Choose the Atom in Each Pair with the Higher First Ionization Energy

1. Al or S, Al is further left
2. As or Sb, Sb is further down
3. N or Si, Si is further down and left
4. O or Cl, opposing trends



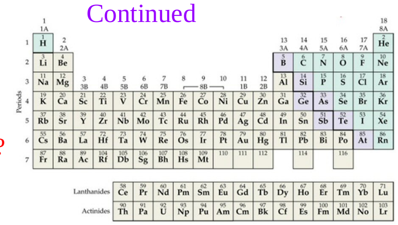
Practice—Choose the Atom with the Highest Ionization Energy in Each Pair

1. Mg or P
2. Cl or Br
3. Se or Sb
4. P or Se



Practice—Choose the Atom with the Highest Ionization Energy in Each Pair, Continued

1. Mg or P
2. Cl or Br
3. Se or Sb
4. P or Se ?



Periodic Table of Ionization Energy

First Ionization Energies of Some Elements (kJ/mol)

Period	Group 1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	H 1310																	He 2370
2	Li 520	Be 900											B 800	C 1090	N 1400	O 1310	F 1680	Ne 2080
3	Na 490	Mg 730											Al 580	Si 780	P 1060	S 1000	Cl 1250	Ar 1520
4	K 420	Ca 590	Sc 630	Ti 660	V 650	Cr 660	Mn 710	Fe 760	Co 760	Ni 730	Cu 740	Zn 910	Ga 580	Ge 780	As 960	Se 950	Br 1140	Kr 1350
5	Rb 400	Sr 550	Y 620	Zr 660	Nb 670	Mo 680	Tc 700	Ru 710	Rh 720	Pd 800	Ag 730	Cd 870	In 560	Sn 700	Sb 830	Te 870	I 1010	Xe 1170
6	Cs 380	Ba 500	La 540	Hf 700	Ta 760	W 770	Re 760	Os 840	Ir 890	Pt 870	Au 890	Hg 1000	Tl 590	Pb 710	Bi 800	Po 810	At . . .	Rn 1030
7	Fr . . .	Ra 510																

https://openstax.org/books/chemistry-2e/pages/6-5-periodic-variations-in-element-properties

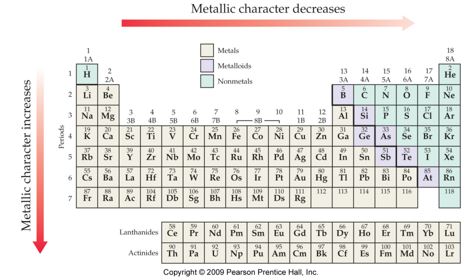
Metallic Character

- How well an element's properties match the general properties of a metal.
- Metals:
  - ✓ Malleable and ductile as solids.
  - ✓ Solids are shiny, lustrous, and reflect light.
  - ✓ Solids conduct heat and electricity.
  - ✓ Most oxides basic and ionic.
  - ✓ Form cations in solution.
  - ✓ Lose electrons in reactions—**oxidized**.
- Nonmetals:
  - ✓ Brittle in solid state.
  - ✓ Solid surface is dull, nonreflective.
  - ✓ Solids are electrical and thermal insulators.
  - ✓ Most oxides are acidic and molecular.
  - ✓ Form anions and polyatomic anions.
  - ✓ Gain electrons in reactions—**reduced**.

Metallic Character, Continued

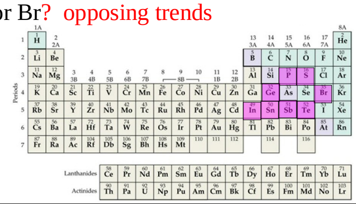
- In general, metals are found on the left of the periodic table and nonmetals on the right.
- As you traverse left to right across the period, the elements become less metallic.
- As you traverse down a column, the elements become more metallic.

Trends in Metallic Character



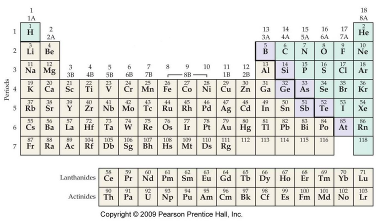
Example—Choose the More Metallic Element in Each Pair

1. Sn or Te, Sn is further left
2. P or Sb, Sb is further down
3. Ge or In, In is further down & left
4. S or Br? opposing trends



Practice—Choose the More Metallic Element in Each Pair

1. Sn or Te
2. Si or Sn
3. Br or Te
4. Se or I



Practice—Choose the More Metallic Element in Each Pair, Continued

1. Sn or Te
2. Si or Sn
3. Br or Te
4. Se or I ?