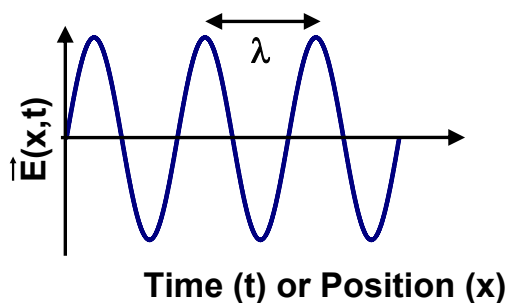


Chapter 7: Atomic Structure and the Periodic Table

Sections 7.1 - 7.2: Light Wavelength and Frequency

Light is an electromagnetic radiation. In space, light travels as a wave. A wave is made of successive crests and troughs. A light wave is described in terms of wavelength and frequency.



Wavelength is the distance between two consecutive crests or troughs. Wavelength is expressed by the symbol λ . λ is expressed in units of length, typically nanometers (nm) or angstroms (\AA)

$$1 \text{ nm} = 10^{-9} \text{ m} \quad \text{and} \quad 1 \text{ \AA} = 10^{-10} \text{ m}$$

Frequency is the number of wave cycles that pass a given point each second. Frequency is expressed by the symbol ν (nu). ν is expressed as inverse seconds (s^{-1}) or Hertz (Hz).

Wavelength (λ) is inversely proportional to frequency (ν).

$$\text{Mathematically, this is expressed as: } \lambda \propto \frac{1}{\nu}$$

Hence, the higher the frequency, the shorter the wavelength of the electromagnetic radiation.

$$\text{or } \lambda \times \nu = c$$

c is the **speed of light** in vacuum ($c = 3.00 \times 10^8 \text{ m.s}^{-1}$)

Note: In vacuum, all electromagnetic radiations travel at a constant speed of $3.00 \times 10^8 \text{ m.s}^{-1}$

$$\text{Remember : } \lambda \times \nu = c \quad \text{or} \quad \lambda = \frac{c}{\nu} \quad \text{or} \quad \nu = \frac{c}{\lambda}$$

Amplitude refers to the intensity of the electromagnetic radiation. The higher the amplitude, the higher the intensity of the electromagnetic radiation.

Example: The sodium vapor lamp often seen on freeways in the US has a wavelength of 587.8 nm. Calculate the frequency of the light wave.

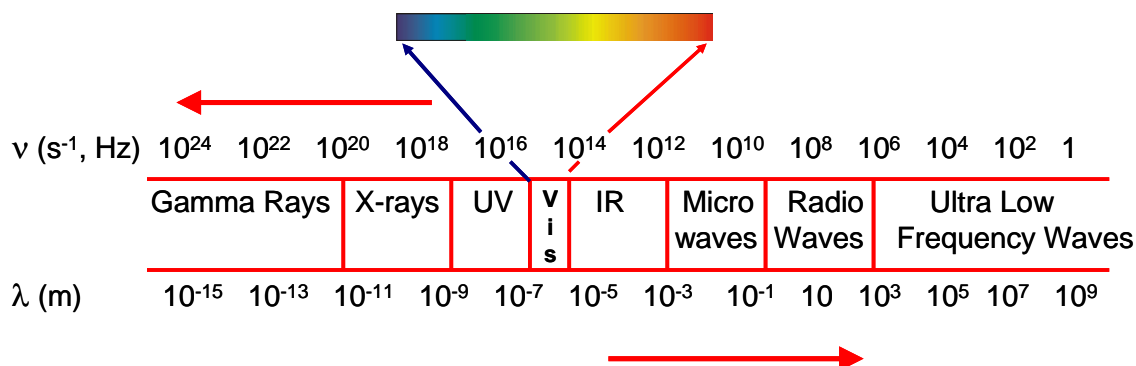
$$\lambda = 587.8 \text{ nm} = 587.8 \text{ nm} \times \frac{10^{-9} \text{ m}}{1 \text{ nm}}$$

$$\nu = \frac{c}{\lambda} = \frac{3.00 \times 10^8 \text{ ms}^{-1}}{5.878 \times 10^{-7} \text{ m}} = 5.1 \times 10^{14} \text{ s}^{-1} \text{ or } 5.1 \times 10^{14} \text{ Hz.}$$

In Section 7.2, practice the Interactive Problems.

Sections 7.3 - 7.4: Electromagnetic Spectrum and Planck's Equation

There are various types of electromagnetic radiation.



In the figure, various types of electromagnetic radiation are arranged in order of increasing wavelength (λ).

Electromagnetic waves can have wavelengths ranging from 10^{-15} m (γ -rays) to 10^9 m (ultra low frequency waves). The whole set of electromagnetic waves (gamma rays, X-rays, ultraviolet (UV), visible, infrared (IR), microwaves, radiowaves and ultra low frequency waves) makes up the **electromagnetic spectrum**.

The electromagnetic radiation consists of tiny particles called **photons**. Electromagnetic radiation can be described as a wave or as a particle (the photon is the smallest packet of electromagnetic radiation).

Energy of light is **quantized** (i.e. a particular electromagnetic radiation contains a certain number of photons). The energy of a light wave is a multiple of the energy of a photon.

The energy of a photon, ε , is proportional to the frequency ν , of the electromagnetic radiation.

$$\varepsilon \propto \nu$$

$$\varepsilon = h\nu = h \frac{c}{\lambda} \quad \leftarrow \text{energy of 1 photon of light having frequency, } \nu, \text{ and wavelength, } \lambda.$$

where h is **Planck's constant** ($h = 6.626 \times 10^{-34}$ J.s).

$$\varepsilon = h\nu \quad \text{is called } \mathbf{Planck's\ equation}$$

The smaller λ or the larger ν , the higher the energy of the photon, ε .

In the visible region, light is broken down into colors. λ ranges from 400 nm – 700 nm. Each color is associated to a characteristic wavelength. When a radiation is constituted with electromagnetic waves characterized by a single wavelength, the radiation is said to be **monochromatic** (single color or single wavelength).

The scattering of light is inversely proportional to the fourth power of λ . Hence, green light scatters more than red light. The more light can be scattered, the less it is visible from a longer distance. For this reason red light is to STOP and green light is to GO.

Example: Calculate the energy in joules for a red light having a wavelength of 629.7 nm.

$$\lambda = 629.7 \text{ nm}$$

$$\lambda = 629.7 \text{ nm} \times \frac{1 \text{ m}}{10^9 \text{ nm}} = 6.297 \times 10^{-7} \text{ m}$$

$$\varepsilon = \frac{6.626 \times 10^{-34} \text{ J.s} \times 3.00 \times 10^8 \text{ m/s}}{6.297 \times 10^{-7} \text{ m}}$$

$$\varepsilon = 3.16 \times 10^{-19} \text{ J}$$

3.16×10^{-19} J is the energy of one photon.

Calculate the energy in kJ of a mole of these photons.

Recall Avogadro's number 1 mol = 6.022×10^{23} photons

$$\varepsilon = 3.16 \times 10^{-19} \text{ J} \times \frac{1 \text{ kJ}}{1000 \text{ J}} \times \frac{6.022 \times 10^{23} \text{ photons}}{1 \text{ mol photons}}$$

$$\varepsilon = 1.90 \times 10^2 \text{ kJ/mol}$$

In Section 7.4, practice the Interactive Problems.

Sections 7.5 - 7.6: Bohr's model of the Hydrogen Atom

In 1913, Niels Bohr proposed a structure for the hydrogen atom. An atom of hydrogen consists of only one electron. Bohr assumed that the electron circles the nucleus on one of many possible orbits and that the electron's energy is **quantized**. Quantized means that the electron energy can have only certain specified values, one for each orbit.

This quantized electron energy is expressed by the symbol E_n .

n – describes the orbit number

n has integral values, which means $n = 1, 2, 3, \dots$ etc.

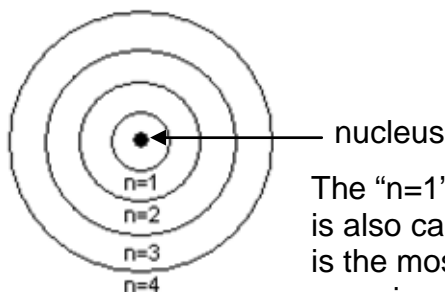
The larger the n value, the further away the orbit is from the nucleus.

The allowable energy values would be E_1, E_2, E_3, \dots and are given by the equation:

$$E_n = \frac{-R_H}{n^2}$$

R_H is Rydberg's constant. $R_H = 2.180 \times 10^{-18} \text{ J}$

This means that an electron in any allowed energy state within the atom must have an energy less than zero. The lowest energy state for the hydrogen atom (orbit $n = 1$) is called the "**ground state**". Negative energies indicate stable states. As the electron moves infinitely far from its nucleus, its energy approaches zero.



The " $n=1$ " orbit (lowest and most negative energy) is also called the **ground state**. The ground state is the most stable state. Other orbits have higher energies and are called **excited states**.

Bohr's model also explains that energy must be absorbed by the atom for an electron to jump from one orbit to another one that is further away from the nucleus (higher n value). We say that the electron makes a transition from one energy level to higher energy level. Atoms can absorb energy either from an electromagnetic radiation of the appropriate energy or when they are heated.

The spectrum generated is called an “**absorption spectrum**”.

When the electron finds itself in an excited state (on an orbit with $n > 1$), it strives to fall back into the ground state. Energy is never lost or destroyed, so when an electron goes from a high energy excited state to the lower energy ground state, it must release some energy. Energy is released through the emission of a photon of light.

This is why such a process is called “**emission**”. Light is emitted with a characteristic frequency, ν , such that the energy of the photon ($\varepsilon = h\nu$) is equal to the difference in energy (ΔE) between the two orbits (starting and ending) for the electron.

In absorption as well as emission, the energy, ε , of a photon involved is ΔE

$$\Delta E = E_{\text{high}} - E_{\text{low}} = h\nu$$

$$E_{\text{high}} = \frac{-R_H}{(n_{\text{high}})^2} \text{ and } E_{\text{low}} = \frac{-R_H}{(n_{\text{low}})^2}$$

$$\Delta E = \frac{R_H}{(n_{\text{low}})^2} - \frac{R_H}{(n_{\text{high}})^2} = h\nu$$

$$\nu = \frac{R_H}{h} \left[\frac{1}{(n_{\text{low}})^2} - \frac{1}{(n_{\text{high}})^2} \right] \quad \text{Recall } \nu = \frac{c}{\lambda}$$

Hence, these equations can be used to calculate the frequency or the wavelength for the transitions that are depicted as lines in the hydrogen spectrum.

Example 1: Calculate the energy of an electron in the third energy level of the hydrogen atom.

$$n = 3 \quad E_3 = \frac{-2.180 \times 10^{-18} \text{ J}}{(3)^2}$$

$$E_3 = -2.422 \times 10^{-19} \text{ J}$$

Example 2: Calculate the frequency of light emitted when an electron falls from $n = 5$ to $n = 3$ in a hydrogen atom.

$$\nu = \frac{2.180 \times 10^{-18} \text{ J}}{6.626 \times 10^{-34} \text{ J}\cdot\text{s}} \left[\frac{1}{(3)^2} - \frac{1}{(5)^2} \right]$$

$$\nu = (3.290 \times 10^{15} \text{ s}^{-1}) \times (0.07)$$

$$\nu = 2.303 \times 10^{14} \text{ s}^{-1}$$

In Section 7.6, practice the Interactive Problems.

Section 7.7: Heisenberg Uncertainty Principle

While Bohr's model was successful at explaining why certain radiations can be absorbed and emitted by atoms, it was physically unrealistic because electrons in such a model should lose energy with time and eventually collide with the nucleus. This does not happen.

Heisenberg, and simultaneously, Schrödinger, thought that it would be better to describe electrons as waves. In the same way that we describe light waves as particles (photons), we can describe particles (electrons) as waves.

Actually, De Broglie proposed that any object of mass (m) and velocity (v) can be described as a wave of wavelength (λ) where

$$\lambda = \frac{h}{mv}$$

← Planck's constant
← Particle velocity
← Particle mass
← Particle wavelength

DeBroglie Equation

If a particle is described as a wave, then, there is a limitation as to the precision with which one can know its position and velocity. Think of the blurry picture one obtains when taking the picture of an object in motion; the faster the object moves, the more blurry the object appears.

$$m\Delta x\Delta v \geq \frac{h}{4\pi} \quad \text{This is the Heisenberg Uncertainty Principle}$$

Δx uncertainty on position, x
 Δv uncertainty on velocity, v
 h Planck's constant

The more accurately one knows the velocity, the less accurately one knows the position of the electron and vice-versa.

An electron should therefore be viewed as a wave spread over space, or a **cloud**. One can never know exactly where the electron is. One can only estimate the probability that the electron is at a specific location at a specific time.

Section 7.8: The Quantum Mechanical Model of the Atom

The uncertainty principle leads to the examination of wave motion on the atomic scale. This is quantum mechanics.

Erwin Schrödinger proposed an equation to describe the behavior of the electron

$$H\psi = E\psi \quad \text{Schrödinger equation}$$

where H is the Hamiltonian operator (a mathematical function accounting for the electron's kinetic and potential energies), ψ is called an **orbital** or a **wave function** and E is the energy associated with a particular wave function. Solutions of the Schrödinger equation are sets of (ψ , E) which give the probability (ψ^2) to find the electron at some location and E the energy of the electron in the corresponding orbital.

Note: The word orbital is used to remind us of the orbit in Bohr's model since ψ gives information about the probable location of the electron. Also, as was the case in Bohr's model, each orbit or orbital (or wave function) is characterized by a specific energy. In contrast with Bohr's model, where orbits were characterized by a single number, an orbital is characterized by three numbers, called **quantum numbers**.

For a given energy level the probability can be depicted by an electron density diagram. The electron whirls around the nucleus in a wave-like motion which appears as a cloud.

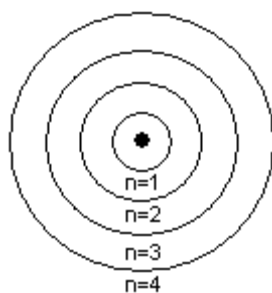
Hence, the electron density diagram is often called an **electron cloud**.

Section 7.9: Quantum Numbers: First Quantum Number or the Principal Quantum Number (n)

The Schrödinger equation can be solved exactly for hydrogen atoms and approximately for all other atoms. The solution for a wave function ψ is associated with a set of quantum numbers.

The **principal quantum or first quantum number (n)** can only take integral values greater or equal to 1. n can equal 1, or 2, or 3, or 4, ...or ∞ .

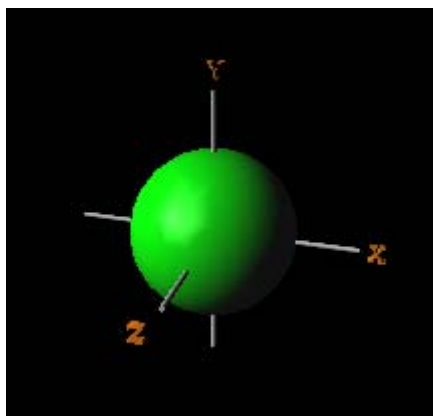
In the case of the hydrogen atom, n fully defines the energy of the electron. (For other atoms, the energy is specified by n and ℓ). n also specifies how far, on average, the electron resides from the nucleus. The larger n, the farther away the orbital containing the electron is from the nucleus. Hence, n defines the size of the orbital. Note that n is also called the **shell number**.



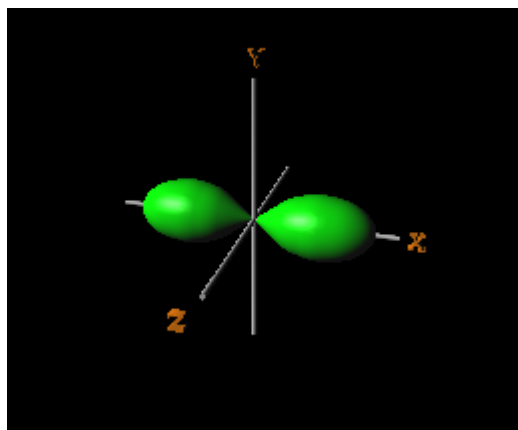
Section 7.10: Second Quantum Number or the Orbital Angular Momentum Quantum Number (ℓ)

ℓ is a non-negative integer that can take any value between 0 and $n-1$. ℓ specifies the shape of an orbital and is also called the subshell number.

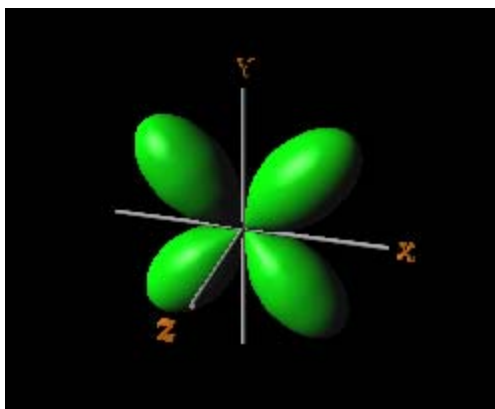
When $\ell = 0$, the orbital is called “s” (spherical shape).



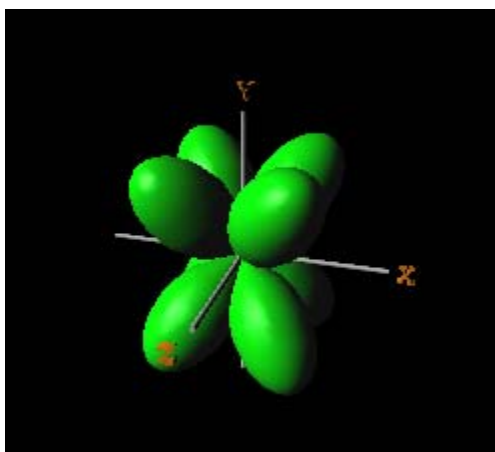
When $\ell = 1$, the orbital is called “p” (dumbbell shape).



When $\ell = 2$, the orbital is called “d”.



When $\ell = 3$, the orbital is called “f”.



For instance, if $n = 1$, the only possible value of ℓ is $\ell = 0$ (since ℓ must have a value between 0 and $n - 1$). Hence, in the shell $n = 1$, one can only find an “s” orbital. This orbital is called “1s.” The “1” in 1s arises from the condition $n = 1$.

If $n = 2$, ℓ can take values of 0 or 1. Possible orbitals are 2s and 2p.

If $n = 3$, ℓ can take values of 0, 1, or 2. Possible orbitals are 3s, 3p, and 3d.

If $n = 4$, ℓ can take values of 0, 1, 2, or 3. Possible orbitals are 4s, 4p, 4d and 4f.

Similarly, one can envision orbitals: 5s, 5p, 5d, and 5f in the shell $n = 5$ or 6s, 6p, 6d and 6f in the shell $n = 6$, etc...

Remember that for a given value of n , the energy of orbitals ns, np, nd and nf are ranked as follows for all elements except H.

$$E_{ns} < E_{np} < E_{nd} < E_{nf} \text{ (ns is the most stable, nf is least stable)}$$

$$\text{For Hydrogen, } E_{ns} = E_{np} = E_{nd} = E_{nf}.$$

Also, as n increases, the energy increases.

$$E_{1s} < E_{2s} \quad E_{2p} < E_{3p} < E_{4p}$$

Section 7.11: Third Quantum Number or the Magnetic Quantum Number (m_ℓ)

m_ℓ , the third quantum number is also called the magnetic quantum number. It is an integer whose value must be between $-\ell$ and $+\ell$. m_ℓ specifies the direction of the orbital.

If $\ell = 0$, m_ℓ must be equal to 0.

There is only one value of m_ℓ for $\ell = 0$ because $\ell = 0$ is a spherical orbital. Two values of m_ℓ would imply two different directions for a sphere. A sphere looks the same in all directions which means that there is only 1 orbital of type s in a given shell.

If $\ell = 1$, m_ℓ can take values of -1, 0, or +1.

Recall $\ell = 1$ is a p-orbital (dumbbell shaped). The dumbbell can be oriented in the x, y or z directions (p_x , p_y , p_z). These orbitals only differ in the direction they point to. They are called **degenerate** orbitals because they have the same energy.

There are 3 "2p" orbitals $2p_x$, $2p_y$, $2p_z$.

There are 3 "3p" orbitals $3p_x$, $3p_y$, $3p_z$.

There are 3 "4p" orbitals $4p_x$, $4p_y$, $4p_z$.

etc.

If $\ell = 2$, m_ℓ can take values of -2, -1, 0, +1, or +2.

There are 5 orbitals of type d in a given shell (given n value). These orbitals are named d_{xy} , d_{xz} , d_{yz} , d_{z^2} , $d_{x^2-y^2}$. Again, these orbitals point in different directions, but have the same energy.

If $\ell = 3$, m_ℓ can take values of -3, -2, -1, 0, +1, +2, or +3.

There are 7 orbitals of type f in a given shell (given n value). These orbitals point in different directions and have the same energy.

In summary, think of a specific orbital (or specific wave function ψ) as the house for an electron. The house is characterized by an address (three quantum numbers: n, ℓ and m_ℓ). n defines the city (shell) where the house is located, ℓ defines the street (subshell s, p, d and f) on which the house is found and m_ℓ defines the street number for the house.

Two houses with same n and ℓ values, but different m_ℓ are two houses (orbitals) in the same street (same subshell " ℓ ") and same city (same shell "n"), but at different numbers.

To see the pictures of the orbitals in different directions, refer to the DVD.

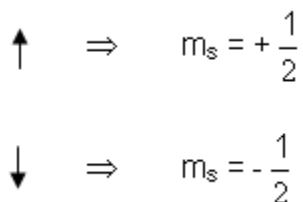
Section 7.12: Fourth Quantum Number or the Electron-Spin Magnetic Quantum Number (m_s)

While n , ℓ and m_ℓ characterize the orbital, the last quantum number characterizes the direction in which the electron is spinning around its own axis.

$m_s = +\frac{1}{2}$ electron is spinning clockwise

$m_s = -\frac{1}{2}$ electron is spinning counterclockwise

Two electrons with the same m_s values are said to have parallel spins. Two electrons with opposite m_s values are said to have anti-parallel spins, or opposite spins. The spin of an electron is often represented by an arrow.



Section 7.13: Pauli Exclusion Principle and Problems on Quantum Numbers

In 1925, Wolfgang Pauli introduced an important principle, "No two electrons in an atom can have the same set of four quantum numbers." This principle has the following implications:

- An orbital can fit a maximum of two electrons. When two electrons occupy an orbital, they must have opposite spins.
- According to this principle, the maximum number of electrons in the s-orbital is 2, the maximum number of electrons in the three p-orbitals is 6, the maximum number of electrons in the five d-orbitals is 10, the maximum number of electrons in the seven f-orbitals is 14.
- Also, for a shell specified by principal quantum number n , the number of orbitals in that shell is n^2 and the maximum number of electrons in that shell is $2n^2$.

For $n = 1$, only 1 orbital (1s) \Rightarrow 2 electrons max.

For $n = 2$, 4 orbitals (2s, 2p_x, 2p_y, 2p_z) \Rightarrow 8 electrons max.

For $n = 3$, 9 orbitals (3s, 3p_x, 3p_y, 3p_z, 3d_{xy}, 3d_{xz}, 3d_{yz}, 3d_z², 3d_{x²-y²}) \Rightarrow 18 e⁻ max.

For $n = 4$, 16 orbitals (one 4s, three 4p, five 4d, seven 4f) $\Rightarrow 32 e^-$ max.

In this section, an interactive problem is given to review the above concepts.

Sections 7.15 - 7.16: Atomic Electron Configuration: The Aufbau Principle

The **atomic electron configuration** is the arrangement of electrons in an atom. In terms of energy, the most stable state is the ground state. The electron configuration of an atom is the arrangement of electrons in the lowest possible energy state.

The Pauli Exclusion Principle lets us fill the orbitals with electrons in order of increasing energy.

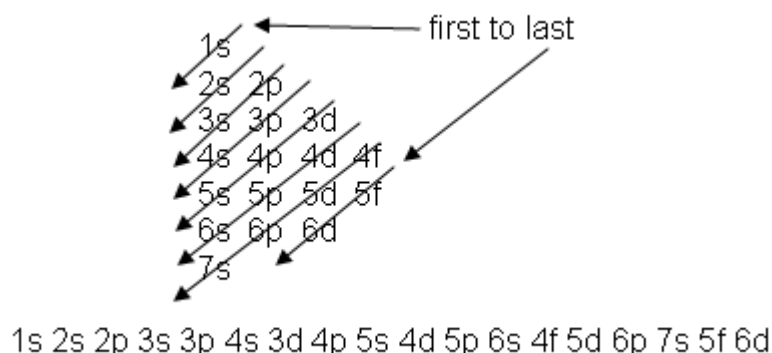
$$ns < np < nd < nf$$

The process of placing electrons in the orbitals (i.e. figuring out which orbitals are filled partially or fully with electrons) is called the “**Aufbau**” or the Build-up process.

Strategy:

We fill the orbitals with electrons in such a way that the lowest energy orbitals are filled first and the least stable (highest energy) orbitals are filled last. We follow Pauli's exclusion principle, meaning that we put no more than 2 electrons in each orbital.

The order in which the orbitals are filled is shown on the schematic below.

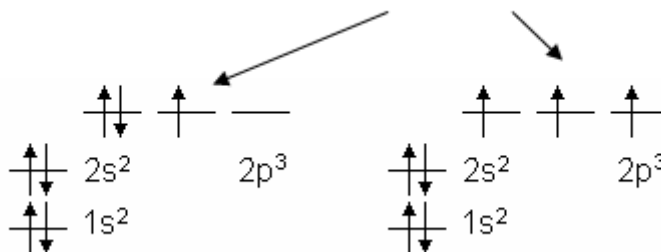


Also, remember that a given subshell (fixed n and l) may have more than one orbital (a “p” subshell ($l = 1$) contains 3 orbitals, a “d” subshell ($l = 2$) contains 5 orbitals, an “f” subshell ($l = 3$) contains 7 orbitals).

This information, as well as the order in which these subshells are filled, can be found on the periodic table itself. Remember that each orbital contains a maximum of two electrons.

In Section 7.16, see the animation of the electron configurations of various atoms. Hund's rule is also explained along with the animations.

Remember: **Hund's rule** states that one should always try to half-fill degenerate orbitals before we fill them completely. This explains why of the two electron configurations shown below for nitrogen (N), only the one on the right is correct.



Section 7.17: Important Trends in Atomic Electron Configurations

The atoms of elements in a group have the same distribution of electrons in the outermost principal energy level.

Elements in **groups 1 and 2 fill the s-sublevel**. For example, the electron configuration of Na is $1s^2 2s^2 2p^6 3s^1$ and the electron configuration of Mg is $1s^2 2s^2 2p^6 3s^2$. Both these elements are filling the s-sublevel.

Elements in **groups 13 through 18 fill the p-sublevel** except He. For example, the electron configuration of Al is $1s^2 2s^2 2p^6 3s^2 3p^1$ and the electron configuration of Ar is $1s^2 2s^2 2p^6 3s^2 3p^6$. Both these elements are filling the p-sublevel.

Elements in **groups 3 through 12 fill d-sublevel**. For example, the electron configuration of Sc is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^1$ and the electron configuration of V is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^3$. Both these elements are filling the d-sublevel.

The set of elements listed separately **at the bottom of the table fill the f-sublevel**. Lanthanides fill the 4f-orbitals. Actinides fill the 5f-orbitals.

The explanation given below is not in the DVD.

Atoms of elements in a group (with some exceptions for groups 3-12) have the same distribution of electrons in their outermost shell (or outermost principal energy level).

For example:

Li, Na, and K in Group 1 are all characterized by a ns^1 outermost shell.

Li	$3e^-$	$2s^1$
Na	$11e^-$	$3s^1$
K	$19e^-$	$4s^1$

Because these elements are in the same group, they have the same electron configuration in the outer shell and they have related chemical properties. The same can be said about F, Cl, Br and I. Recall the reactions of Li, Na, and K with H₂O shown in the Chapter 1 videos on your DVD.

Section 7.18: Abbreviated Atomic Electron Configurations

Since Noble elements have completely filled orbitals, they are used in expressing atomic electron configurations of other elements.

K has 19 e⁻ ⇒ Electron configuration is: 1s²2s²2p⁶3s²3p⁶4s¹

The noble element just before K is Ar. The electron configuration of Ar is: 1s²2s²2p⁶3s²3p⁶. Hence, we write the electron configuration of K as: [Ar] 4s¹ where the [] indicate the filled shell electron configuration of a noble element.

[Ar] 4s¹ is the **abbreviated electron configuration** of K.

Example: Write the abbreviated electron configuration of vanadium.

V has 23 e⁻ ⇒ 1s²2s²2p⁶3s²3p⁶4s²3d³

⇒ V = [Ar] 4s²3d³

Exceptions to Aufbau for electron configurations of transition metals:

Cu has 29e⁻, so its predicted configuration is: [Ar] 4s²3d⁹
However, Cu actually has the following configuration ⇒ [Ar] 4s¹3d¹⁰

Cr has 24e⁻, so its predicted configuration is: [Ar] 4s²3d⁴
However, Cr actually has the following configuration ⇒ [Ar] 4s¹3d⁵

Some of the other exceptions are: Nb, Mo, Ru, Rh, Pd, and Ag in Period 5 and Pt and Au in period 6. Understanding the origin of these exceptions is beyond the scope of freshman chemistry and requires a deeper understanding of quantum mechanical calculations.

Section 7.19: Electron Configuration of Ions

Elements close to noble gases form ions that have the same number of electrons as the atom of a noble gas. A nitrogen atom would gain 3 electrons to become N⁻³ and N⁻³ has the same number of electrons (10 e⁻) as Ne. We say that N⁻³ and Ne are **isoelectronic** (that is they have the same number of electrons and the same electron configuration).

S⁻² is isoelectronic with Ar.

Br^- is isoelectronic with Kr.

For metal ions, one must differentiate metals from the s-block from metals in the p-block or metals from the transition d-block.

1. Group 1 and 2 metals form ions (cations) which also exhibit the electronic configuration of the nearest noble element.

Na loses $1e^-$ to form Na^+ which has the same electron configuration as Ne.

Li loses $1e^-$ to form Li^+ which has the same electron configuration as He.

Ca loses $2e^-$ to form Ca^{+2} which has the same electron configuration as Ar.

2. Group 3 through 12 transition metals can lose a variable number of electrons. For instance, Fe can lose 2 or 3 e^- to form Fe^{+2} and Fe^{+3} cations. Cu can lose 1 or 2 e^- to form Cu^+ and Cu^{+2} . In the case of transition metals, the “s” electrons are removed before the “d” electrons.
3. Group 13-17 elements, when forming cations, lose p-electrons before losing s-electrons.

In Section 7.21, practice the atomic electron configuration.

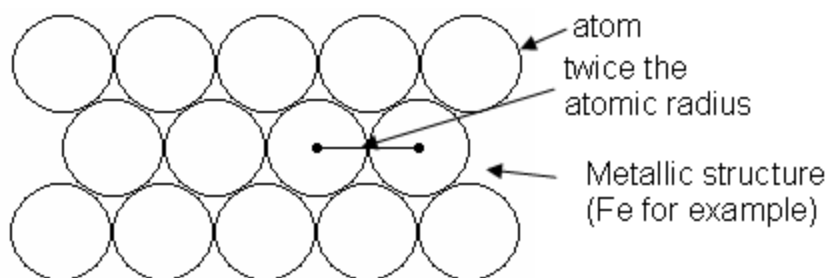
Section 7.22: Periodic Trends in Some Key Atomic Properties: Atomic Size

The chemical and some physical properties of elements are based on atomic electron configurations. The three atomic properties that are directly influenced by the electron configuration are:

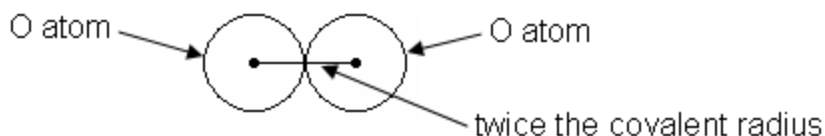
Atomic size
Ionic size
Electron affinity
Ionization energy

1. Atomic Size

For metallic elements, the **atomic size** (also known as the **metallic radius**) is defined as half the **interatomic distance** in the solid metal.



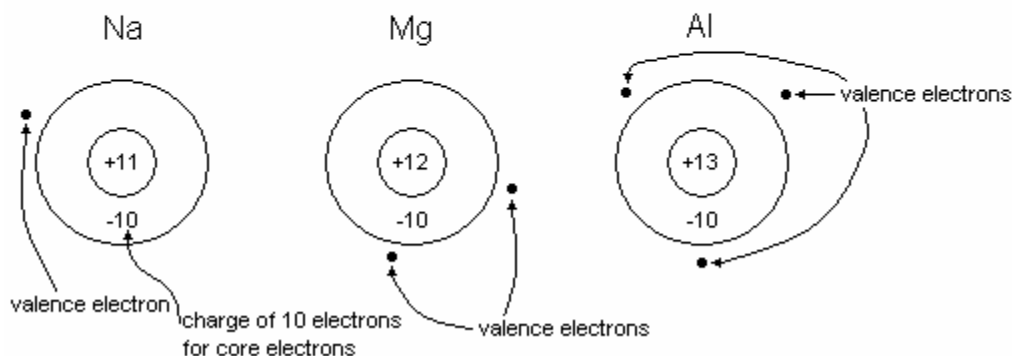
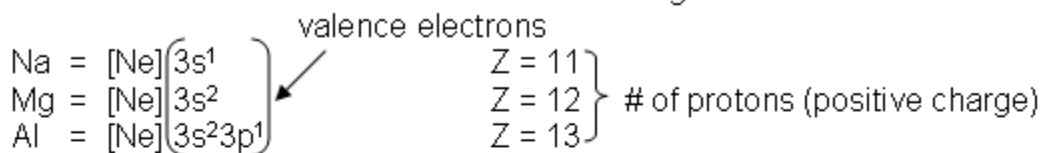
For non-metal elements, the atomic size (also known as **covalent radius**) is defined as half the **bond length** in the molecular form of the element.



In a group, the atomic size increases from top to bottom. This increase in atomic size is explained by the increasing number of filled shells as one goes down in a group.

In a period, the atomic size decreases from left to right. The decrease in atomic size is explained by an increase in the nuclear charge as one travels from left to right in a period. To understand the role of the nuclear charge, consider Na, Mg and Al (3 elements of the same period ($n=3$)).

All three elements have the core electron configuration of Neon.



To decide which of these atoms has the smallest radius, we consider the strength of attraction between each of the valence electrons and the nucleus. We should indeed

expect that the stronger the attraction between the valence electrons and the nucleus, the closer the valence electrons to the nucleus and the smaller the atom.

A valence electron is attracted by the nucleus since they have opposite charges and is repelled by the core electrons (charges of the same sign). Hence, the actual attraction of a valence electron by the nucleus depends on the charge of the nucleus and on the number of core electrons between the nucleus and the valence electrons. We say that the core electrons **shield** the valence electrons from the nucleus and decrease the nuclear charge felt by the valence electron. The actual nuclear charge “felt” by the valence electrons is called the **effective nuclear charge** and denoted Z_{eff} .

Z_{eff} is always less than the true nuclear charge Z . The larger the number of core electrons the larger the difference between Z and Z_{eff} .

We note that Na, Mg and Al have the same number of core electrons ($10e^-$) since they belong to the same period. The true nuclear charge ($+Z$) increases from 11 for Na to 12 for Mg and to 13 for Al. Hence, the effective nuclear charge increases from Na to Mg to Al. Thus, the attraction between a valence electron (outer e^-) and the nucleus increases from Na to Mg to Al. As a result of this increased attraction, the size of the atom decreases in the order $\text{Na} > \text{Mg} > \text{Al}$.

2. Ionic Size

A cation is smaller than the parent atom.

An anion is larger than the parent atom.

If $\text{Li} > \text{Be} > \text{B} > \text{C} > \text{N} > \text{O}$

Then, $\text{Li}^+ > \text{Be}^+ > \text{B}^+ > \text{C}^+ > \text{N}^+ > \text{O}^+$

Such trends are rather useless to know, since B^+ , C^+ , N^+ and O^+ are rather rarely encountered. A more useful trend is that of a series of isoelectronic species.

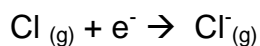
$\text{F}^- > \text{Ne} > \text{Na}^+ > \text{Mg}^{+2} > \text{Al}^{+3}$

These 5 species have the same number of electrons ($10e^-$). Hence, they have the same e^- configuration $1s^22s^22p^6$. F^- is larger than Ne , Na^+ , Al^{+3} because F^- has a smaller effective nuclear charge than Ne , Na^+ , and Al^{+3} .

The larger the effective nuclear charge, the smaller the atom/ion for a given number of electrons.

Section 7.23: Periodic Trends in Some Key Atomic Properties: Electron Affinity

Electron affinity is the measure of an atom's ability to attract an electron to itself. Electron affinity, E_a , is the energy released during the reaction of an atom in the gaseous state to form an anion.



Note that E_a values are generally negative since for most elements, energy is released during this process. Therefore, the more negative E_a , the more energy is released during the anion formation and the more stable the anion.

Hence, the most negative values are found for the electron affinity of elements in groups 16 and 17, since these elements are non-metals and like very much to form anions. Noble gases have positive electron affinities since they have filled shells and do not want to gain (or lose) any electrons.

In a group, the magnitude of the electron affinity *generally* decreases from top to bottom, because from top to bottom the atomic radius increases. Similarly, in a period, the magnitude of the electron affinity *generally* increases from left to right because the atomic size decreases from left to right.

Less energy is released when an electron is added far away from the nucleus than when it is added close to the nucleus. Remember that the magnitude of the electrostatic energy is greater when the charges are closer. As inferred from the use of the word *generally* in the above sentences, there are a number of exceptions to this rule.

Section 7.24: Periodic Trends in Some Key Atomic Properties: Ionization Energy

Ionization energy is a measure of an atoms ability to lose an electron. Energy is always required for the removal of an electron from a neutral atom, therefore ionization energies are always positive.

The First Ionization Energy is associated with the removal of the first electron and Second Ionization Energy is associated with the removal of the second electron.

The first ionization energy generally increases from left to right in a period and decreases from top to bottom in a group. The smaller the atom, the more tightly bound the electrons are to the nucleus and the more energy is required to remove them. Alkali metals (group 1) have the lowest first ionization energies while noble gases have the largest values.

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