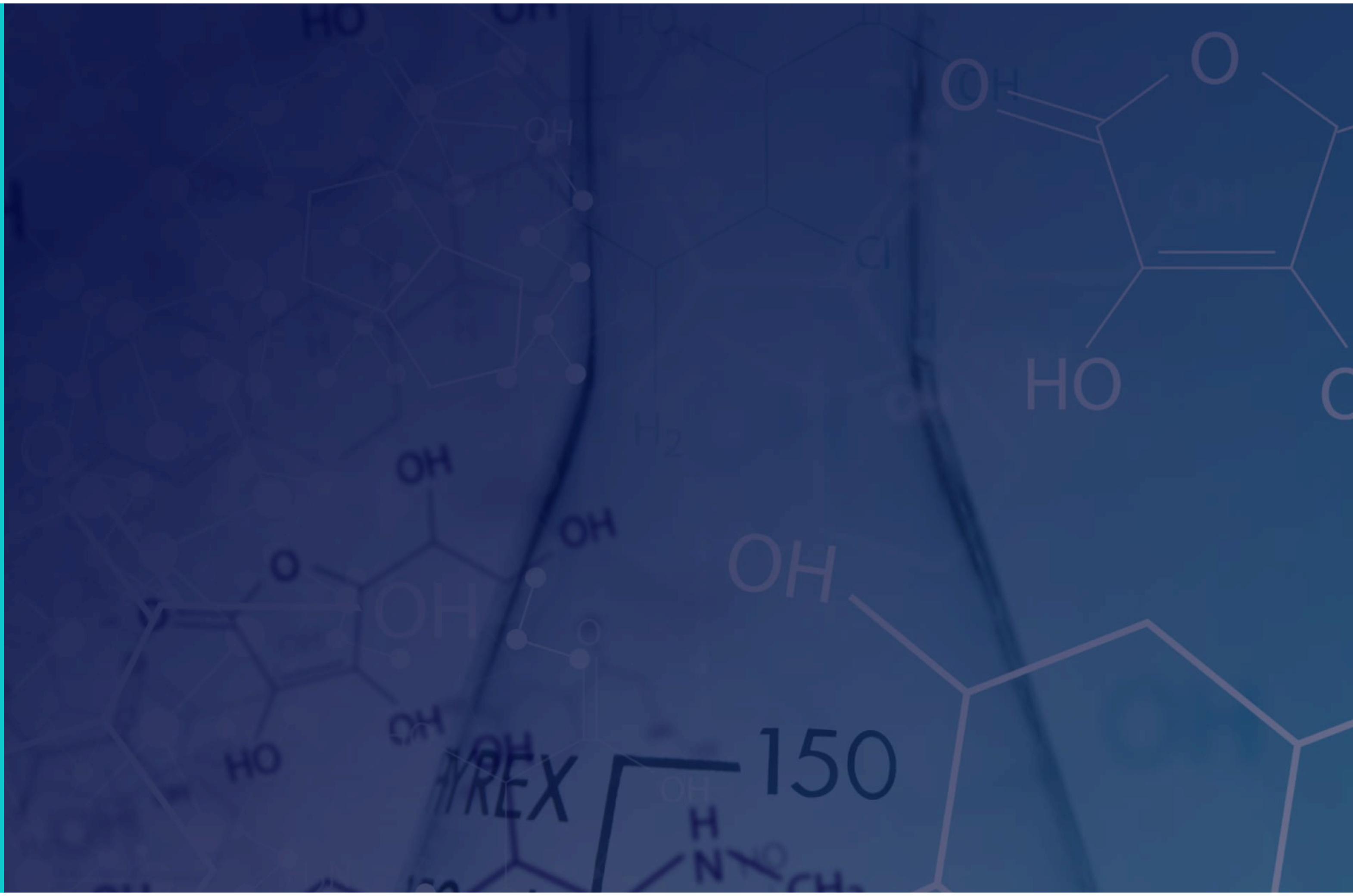

Mac

CHEM 1A03

Fall 2024, Chapter 2 Notes

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2. Unit 3: Atomic Structure and Theory

2.1 Electromagnetic Radiation

2.1.1

Electromagnetic Spectrum

Up until this point in history, scientists were unsure if light shared the properties of waves or particles...there was some evidence for both!

The discovery of the electromagnetic spectrum showed the different wavelengths and frequencies that electromagnetic radiation could exist at. Since the different **types of radiation in this spectrum behave as waves**, and visible light is a part of this spectrum, it became clear to scientists that **light acts as waves!**

Types of Electromagnetic Radiation

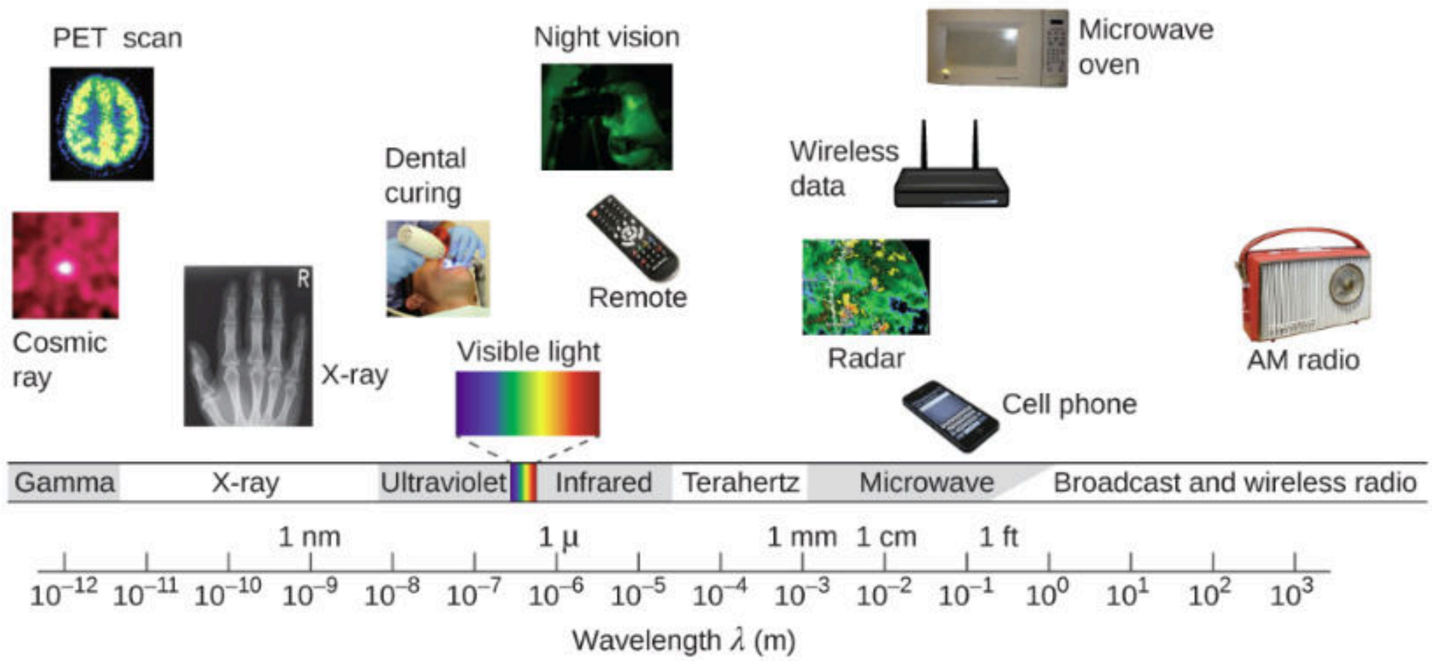


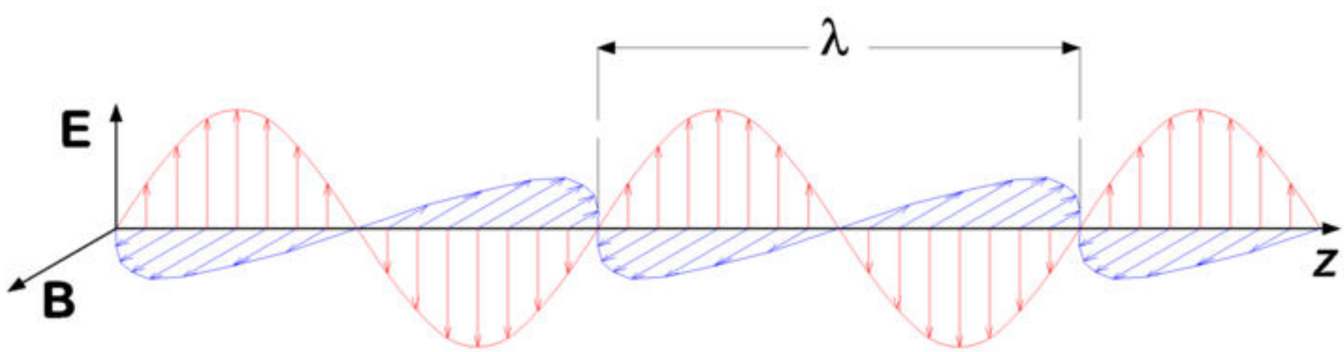
Photo by Rice University/ CC BY

- In the diagram above, the electromagnetic radiation is organized from the types of radiation that are **highest in energy (gamma rays)** to the types of radiation that are **lowest in energy (radiowaves)**
- You might recognize a lot of common household objects on the side of the spectrum with lower amounts of energy!

Photon Equations

Electromagnetic Radiation Waves

- Light (or electromagnetic radiation) behaves as a wave
- A wave is a self-propagating transverse oscillation of electric and magnetic fields



where E is the electric field and B is the magnetic field

- **To describe waves, we can use the following terms:**
  - **Amplitude (A):** the height of a wave crest or depth of a trough
  - **Wavelength (λ):** The distance a wave travels in one cycle **(m)**
  - **Frequency (*v*):** cycles per second (s^{-1} or units Hz)
  - **Energy (E):** units are Joules (J)

Energy of a Photon Equations

$$\boxed{E = \frac{hc}{\lambda} = hv}$$

*where **h is Planck's constant: 6.626x10^{-34} Js***
*and **c is the speed of light: 3.0 x10^{8} m/s***

WIZE TIP

You should **memorize the value for the speed of light (c) = 3.0x10^{8}m/s** as that value is not always provided on exams.

Photon Equations (Cntd.)

Key Relationships Between Variables in Photon Equations

$$E = \frac{hc}{\lambda}$$

- **Are the E and λ of a photon directly or inversely proportional?** ____________ proportional

$$E = hv$$

- **Are the E and *v* of a photon directly or inversely proportional?** __________ proportional

$$\frac{hc}{\lambda} = hv$$

- **Are the wavelength and frequency of a photon directly or inversely proportional?** ____________ proportional

Summary

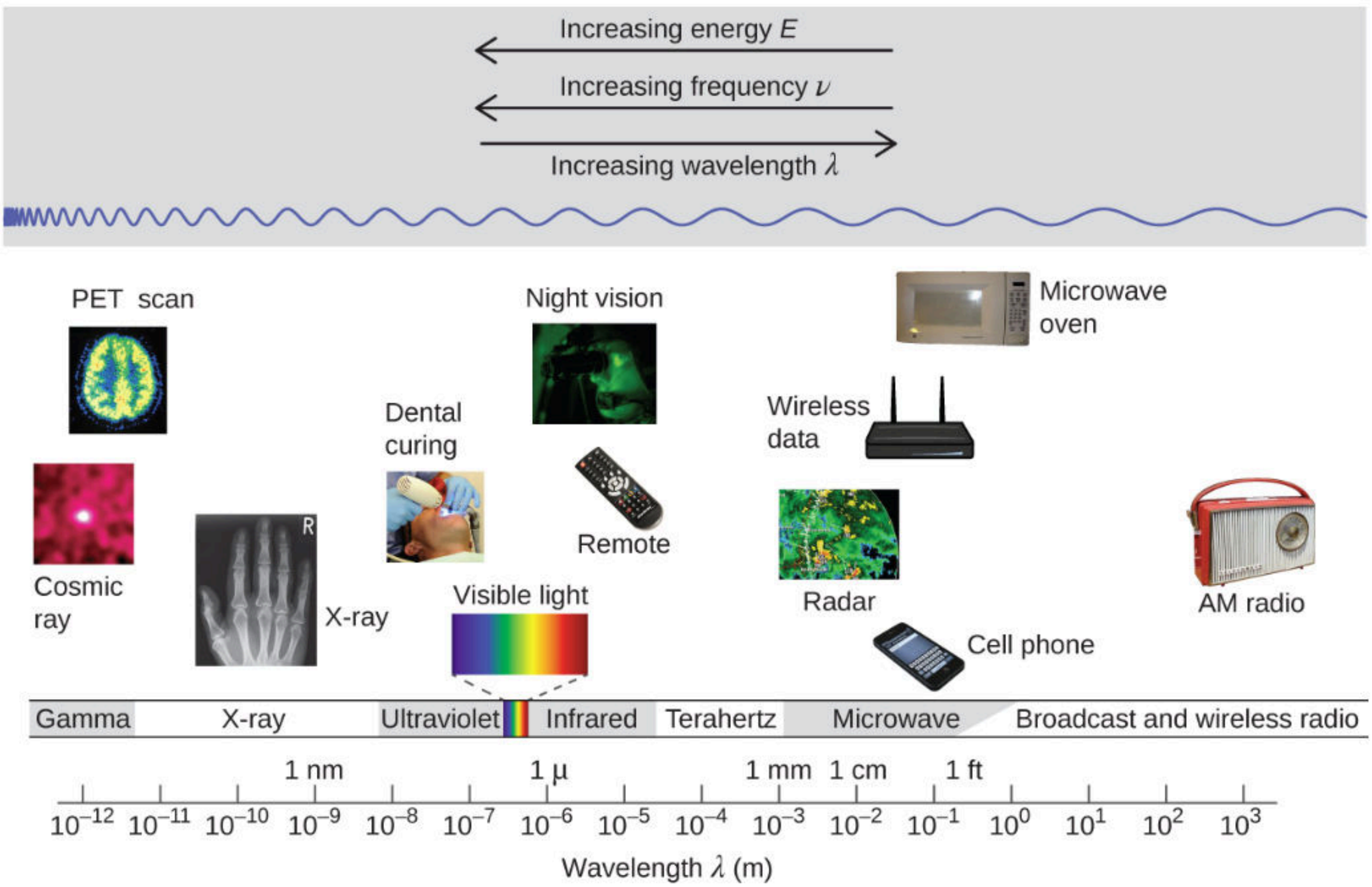


Photo by Rice University/ CC BY

2.1.4 Example

Example: Calculating Energy of a Photon Given Wavelength

A photon of light was found to emit light at 723 nm. How much energy is associated with this light? $h = 6.626 \times 10^{-34}$ Js

2.1.5

Example: Calculating Frequency of a Photon Given Energy

Calculate the frequency of a photon with an energy of 236 kJ.
$h=6.626x10^{-34}$ Js

2.1.6

Practice: Different Types of Electromagnetic Radiation

Rank the following types of electromagnetic radiation from lowest to highest energy.

Radio Waves, Infrared Radiation, X-rays, Microwaves, Visible Light, UV Radiation

- X-rays < Infrared Radiation < Microwaves < UV Radiation < Visible Light < Radio Waves
- Radio Waves < Microwaves < Infrared Radiation < Visible Light <UV Radiation < X-rays
- Radio Waves < Infrared Radiation < Microwaves < Visible Light < UV Radiation < X-rays
- Radio Waves < Infrared Radiation < Visible Light < Microwaves < UV Radiation < X-rays

2.1.7

Practice: Calculating Frequency of a Photon Given Wavelength

A photon of light has a wavelength of 700 nm. What is its frequency?

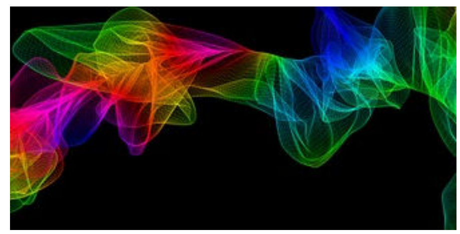
- $4.29x10^{-11}$ s^{-1}
- $6.72x10^{12}$ s^{-1}
- $4.29x10^{14}$ s^{-1}
- $7.23x10^{16}$ s^{-1}

2.1.8

Wave-Particle Duality

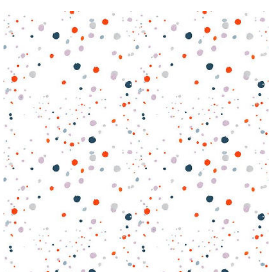
Light was originally thoughy of as only having wave-like properties, but was later shown to have particle-like properties as well.

Light as Waves



- When light was shown at a slit, interference patterns were recorded
- **Interference patterns** are a property of waves

Light as Particles



- Einstein showed that light can also be thought of as particles since they are like packages of exact amounts of energy
- The experiment was the **photoelectric effect** and it showed that there is **wave-particle duality** (light has both properties of waves and particles!).

What About Electrons?

Louis de Broglie thought that if light has both properties of waves and particles, electrons might also have properties of particles and waves.

He determined the following equation to explain the wave-nature of particles:

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

De Broglie Equation

λ *is the* ***de Broglie wavelength*** *of the small particle (****m****)*

h is Planck's constant*= 6.626 x10^{-34} Js (or 6.626x10^{-34} kg m^2/s)*

m *is the mass of the particle (****kg****)*

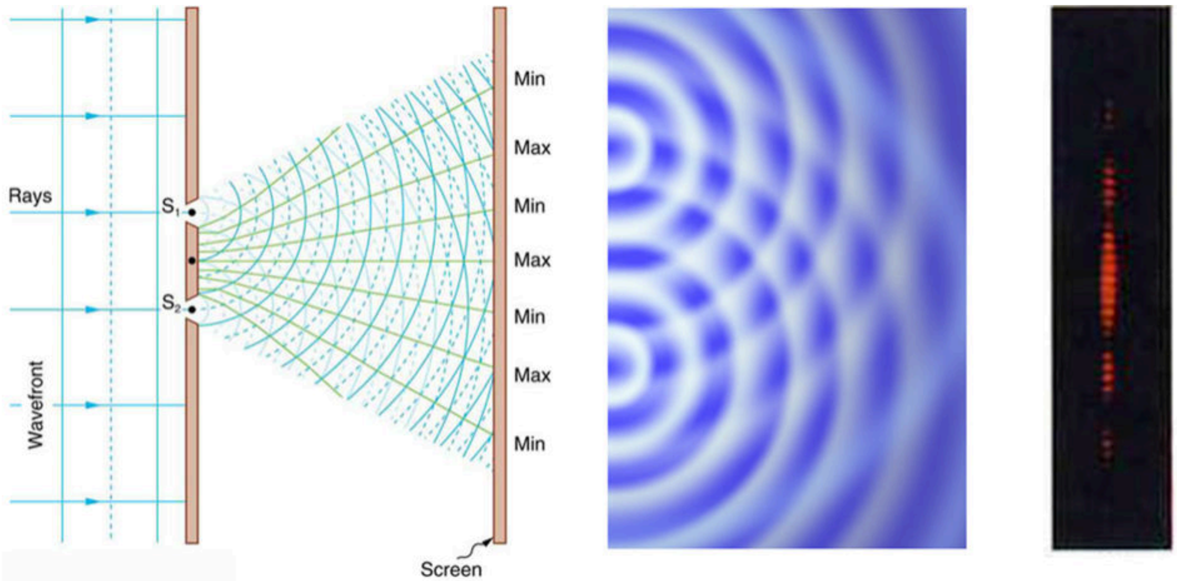
v *is the velocity of the particle (****m/s****)*

p *is the momentum of the particle (p=mv)*

WATCH OUT!

The *v* in this equation is not frequency (*v*, Hz), it is velocity (*v*, m/s).

Light acting as a wave:



Young's double slit light experiment

Could electrons also act as waves?

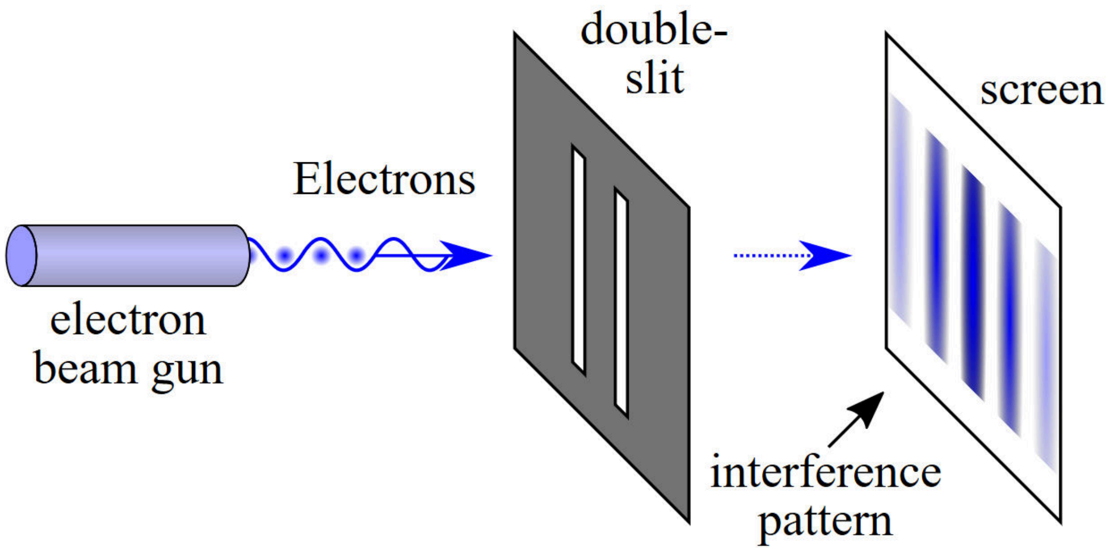


Photo by Johannes Kalliauer / CC BY

In summary

All waves can behave like particles and all particles can behave like waves! **The wave-particle duality applies to light and particles!**

Example: Wavelength of an Accelerated Neutron

At CERN, a neutron is accelerated to 2.4 x 10^{8} m/s in a particle accelerator. The mass of the neutron is 1.675 x 10^{-27} kg, what is it's wavelength? (h is 6.626x10^{-34} kg m^{2}/s)

a) 1.14 x 10^{-15} m
b) 2.77 x 10^{-14} m
c) 1.65 x 10^{-15} m
d) 3.53 x 10^{-14} m

2.1.10

Practice: Two-Slit Experiment

If you shoot a beam of electrons at a piece of metal with two small slits in it, a diffraction pattern appears on the detector on the otherside. This is often referred to as a two slit experiment. This result shows that,

- electrons can behave as particles
- electrons are very small
- electrons can behave as waves
- electrons can behave as both particles and waves
- electrons travel very fast

Practice: Wave-Particle Duality

Nolan Ryan, a fastball pitcher from the 1970’s, was recorded throwing a four seam fastball around 108mph (174km/h). A MLB baseball weighs approximately 145g. What is the wavelength of the baseball?

2.2 Prelude to Quantum Theory

2.2.1

Quantum Numbers

Quantum numbers describe where electrons are positioned around atoms:

Letter	Quantum Number	Description
n	Principal	Size
l	Orbital Angular Momentum	Shape
m_l	Magnetic	Orientation
m_s	Electronic Spin	Electron Up or Down

Rules:

n

- Can be **any positive integer**
  Example: n = (1, 2, 3, 4...)

- As n increases, energy and size of shell ___________

WIZE TIP

n is also called the "principle quantum number"

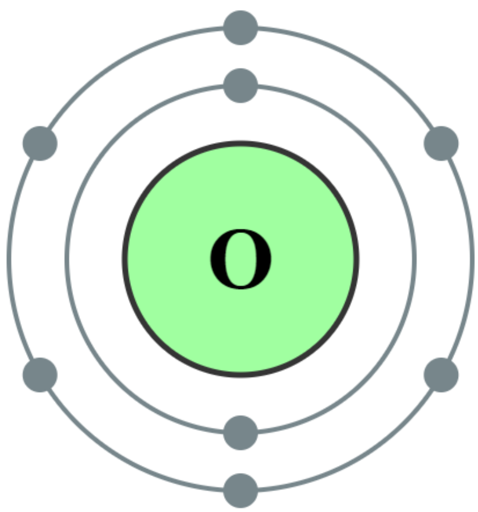


Photo by Greg Robson / CC BY

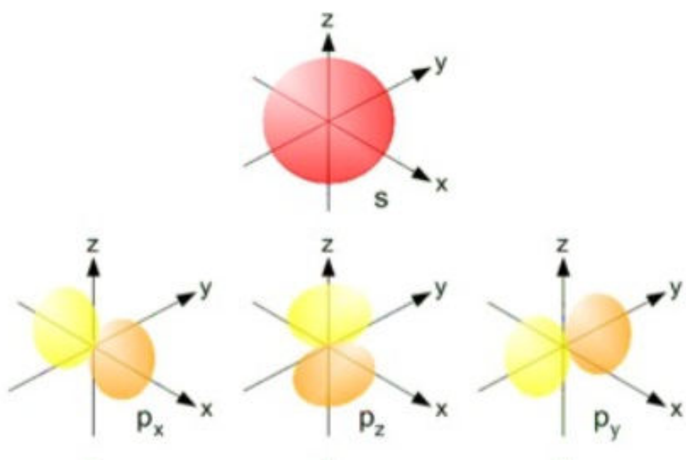
l

- Can be any **non-negative integer up to n-1**
  Example: l = (0, 1, 2, 3, ..., n-1)

- l (orbital shapes) are described as:

subshell	s	p	d	f
l	0	1	2	3

Increasing energy →



- These shapes show us where an electron is most likely to be found
- There is a 90% chance in finding an electron somewhere inside the given shape

WIZE CONCEPT

S orbitals are spheres and with increasing "n" the sphere will get larger.
There is one s orbital in a subshell.

P subshells are dumbbell shaped.
There are three p orbitals in a subshell, each orientated along an axis (px, py and pz)

m_l

- Can be any integer from **–l to +l**
  Example: m_1 = (-l, ..., 0, ..., l)

- This quantum # **designates a specific orbital within a given shell**
  Example: if n=2 and l=1, we are looking at 2p.
  - ml can be -1, 0, or +1 this designates each of the 2p orbitals: 2px, 2py, and 2pz

m_s

- **Spin** of an electron

- Can only be **+1/2 or -1/2**

Quantum Numbers Summary

Principle Quantum Number (n)	Angular Momentum Quantum Number (ℓ)	Magnetic Quantum Number (m_ℓ)	Electron-Spin Quantum Number (m_s)
$n = 1, 2, 3, \ldots$	$\ell = 0, 1, 2, 3, \ldots (n - 1)$	m_ℓ = integer values of ℓ to $-\ell$	$m_s = \pm \frac{1}{2}$
$n \uparrow \rightarrow E \uparrow$	$\ell = 0 \rightarrow s$ $\ell = 3 \rightarrow f$ $\ell = 1 \rightarrow p$ $\ell = 2 \rightarrow d$	**$2\ell + 1$** orbitals	Only 2 electrons per orbital, **ALWAYS**

n shell	ℓ subshell	Orbital	m_ℓ	Number of Orbitals per subshell = n^2
1	0	1s	0	1
2	0	2s	0	4
	1	2p	+1, 0, -1	
3	0	3s	0	9
	1	3p	+1, 0, -1	
	2	3d	+2, +1, 0, -1, -2	
4	0	4s	0	16
	1	4p	+1, 0, -1	
	2	4d	+2, +1, 0, -1, -2	
	3	4f	+3, +2, +1, 0, -1, -2, -3	

2.2.3

Example: Allowed Quantum Numbers

What are the allowed set of quantum numbers for the following orbitals?

a) 6s

b) 4p

c) 3d

Example: Allowed Sets of Quantum Numbers

Which of the following sets of quantum numbers (n, l, m_l, m_s) are allowed and which are not allowed? For the sets of quantum orbitals that are not allowed, state why it is not allowed.

(i) (4, 0, 0, 0)

(ii) (3, 1, 2, $-^1/_2$)

(iii) (5, 3, 0, $+^1/_2$)

(iv) (4, 4, 3, $-^1/_2$)

Example: Defining Orbitals from Quantum Numbers

Determine the atomic orbital described by the following sets of quantum numbers (n, l, m_l, m_s).

(i) (2, 0, 0, $-^1/_2$)

(ii) (4, 3, 0, $+^1/_2$)

(iII) (5, 1, 1, $-^1/_2$)

2.2.6

Practice: Valid Quantum Numbers

Determine which of the following sets of quantum numbers ($n,\ l,\ m_l,\ m_s$) is valid for a 3d orbital.

a) (3, 3, 2, +$^1/_2$)

b) (4, 2, 1, -$^1/_2$)

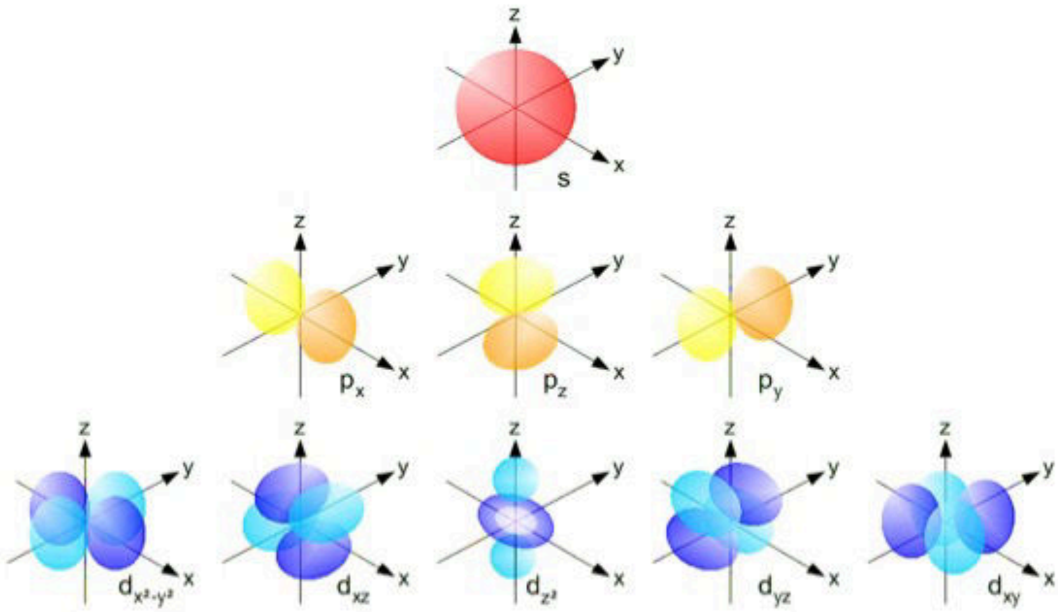
c) (3, 2, -2, 0)

d) (3, 2, -1, -$^1/_2$)

e) (3, 2, 3, -$^1/_2$)

f) (3, 1, 0, +$^1/_2$)

Shapes (aka Boundary Diagrams) of Atomic Orbitals



- There is **one** s orbital in a subshell, which is a sphere
- There are **three** p orbitals in a subshell, each orientated along an axis (p_x, p_y and p_z)

- There are **five** d orbitals in a subshell. The shapes of the d orbitals are more complicated than their s and p counterparts.

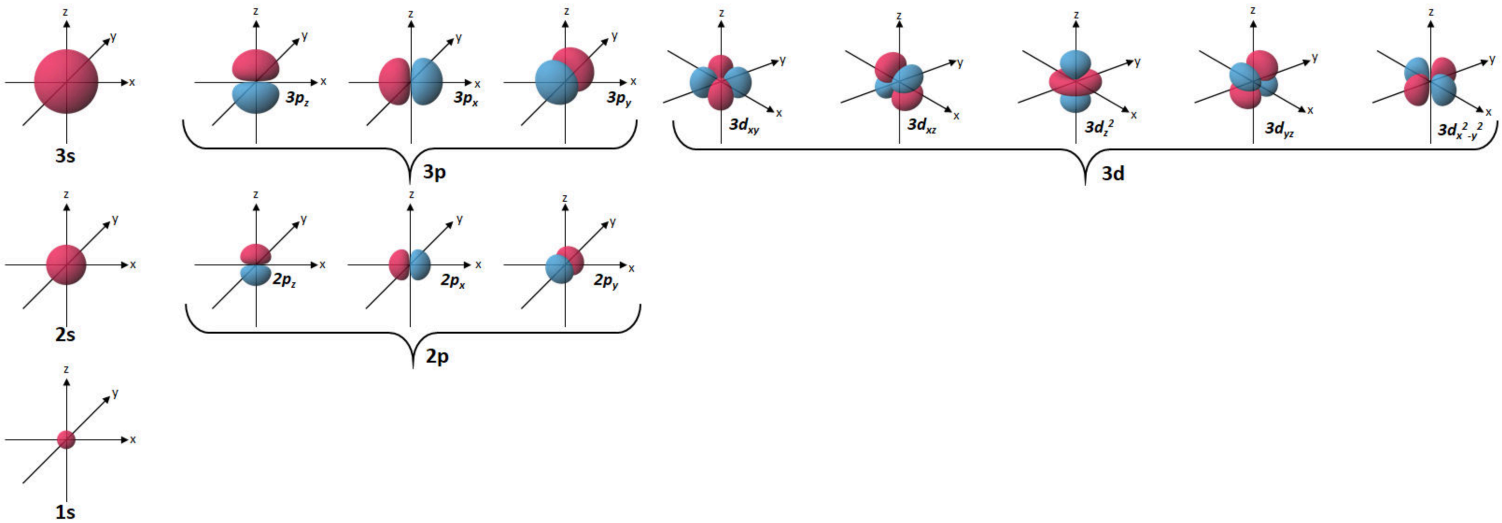
- Three orbitals look like 3D cloverleaves, each lying in a plane with the lobes pointed between the axes (d_{xy}, d_{xz} and d_{yz}).
- A fourth orbital is also a cloverleaf, but its lobes point along the axes ($d_{x^2-y^2}$).
- The fifth orbital looks quite different with its major lobes pointing along the z- axis, but there is also a “doughnut” of electron density in the xy plane (d_{z^2})

Things to Keep in Mind:

- As the **energy level increases**, so does the **size of the orbitals**
  Example: A 2s orbital is higher in energy and larger than a 1s orbital

- As the **energy level increases**, so does the **number of orbitals**
  Example: when n=1 (lower energy level), we see 1 orbital, whereas when n=2 (higher energy level), there are 4 possible orbitals

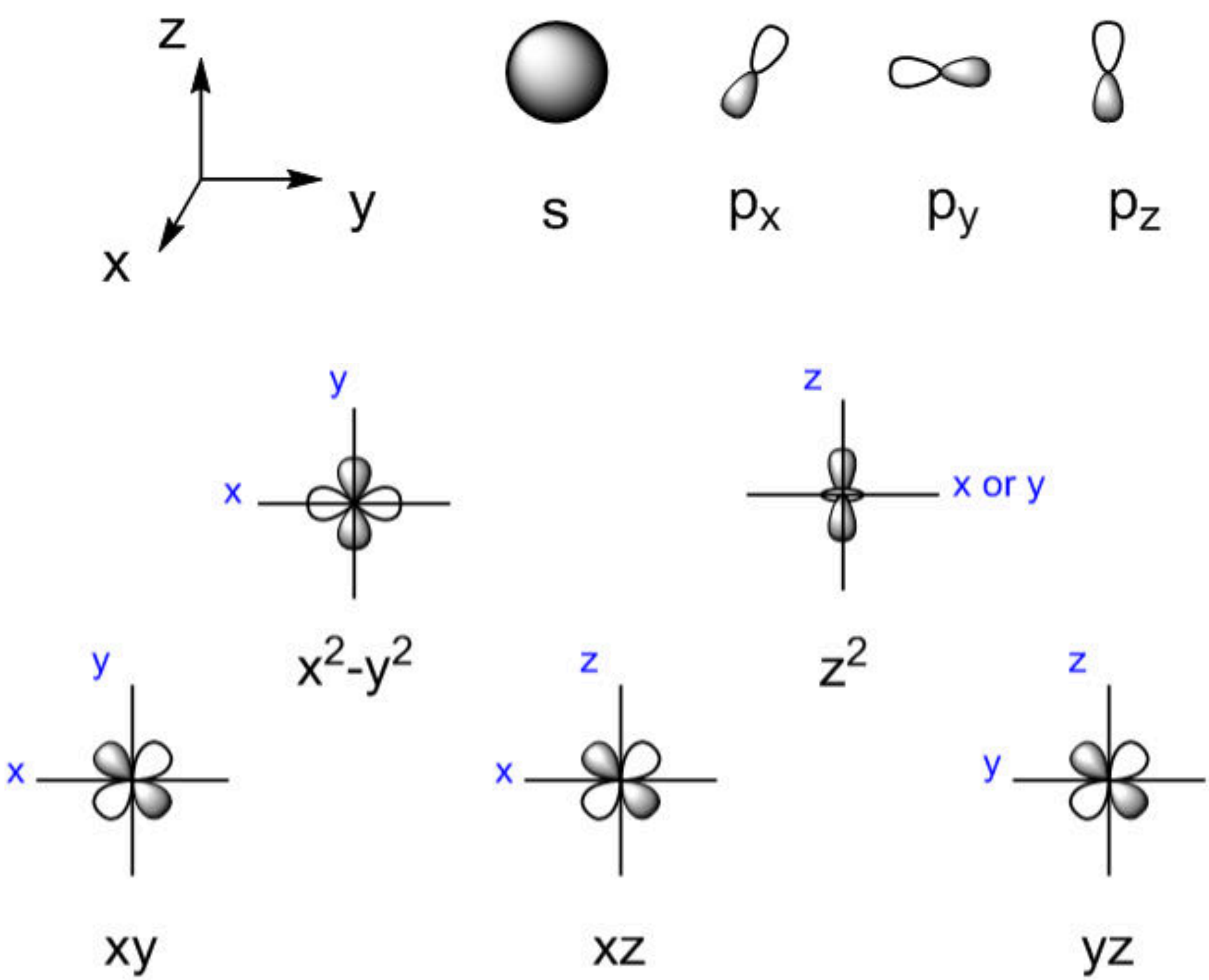
- As the **number of orbitals increases**, so does their **complexity**
  Example: d orbitals are more complex than p orbitals



2.2.8

Drawing Atomic Orbitals

Shown below are the s, p and d orbitals:



WIZE TIP

Pay attention to whether the lobes are ON an axis or BETWEEN axes!

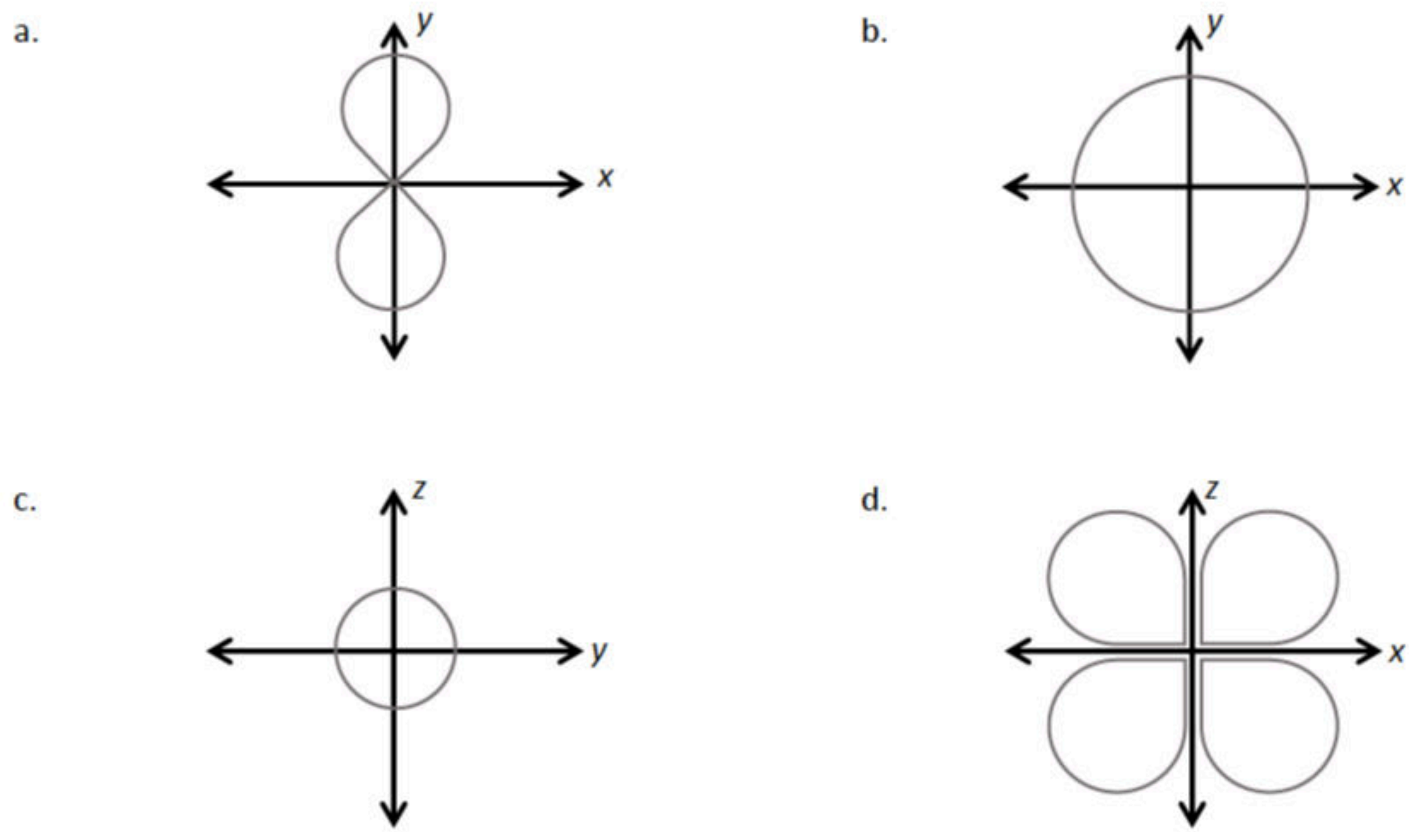
Also, notice how every other lobe is colored in gray. This gray and white coloring is indicating a "positive" lobe and a "negative" lobe. If you were asked to draw one of these orbitals on an exam, you would be expected to color in the lobes as well (or indicate which are + and which are -).

- The angular momentum quantum number, l, gives us information about the shape of the orbital.
- **Note:** There are 5 d-orbitals and 5 allowed value for m_l. Similarly for s and p the number of orbitals is determined by the # of allowed values of m_l.
- **Note:** Each ml is not associated with a particular orbital (ex. d_{xy} vs d_{xz}). Which means given a set of quantum numbers (n,l,ml) one cannot determine which direction the orbital will be pointing

2.2.9

Example: Orbitals

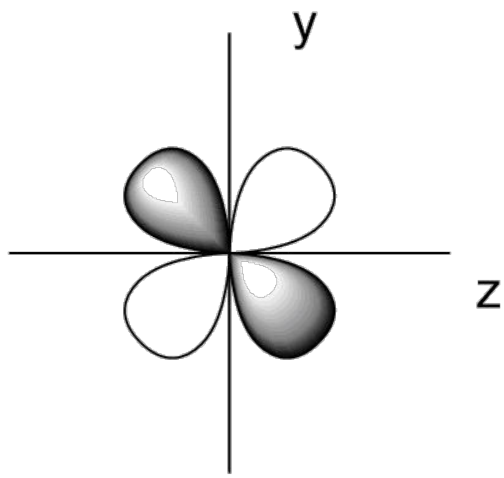
The sketches below show possible orbitals for the electron in a hydrogen atom. Which orbital would have the *lowest* energy?



2.2.10

Practice: Identify the Orbital

Identify the orbital shown below,



a) $2p_y$

b) $3d_{yz}$

c) $3p_x$

d) $3d_{xy}$

e) $3d_{x2-y2}$

Angular and Radial Nodes

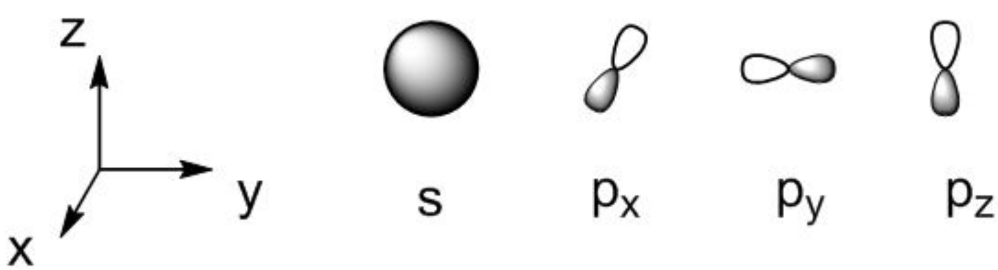
$$\boxed{Total\ number\ of\ nodes\ = n - 1}$$

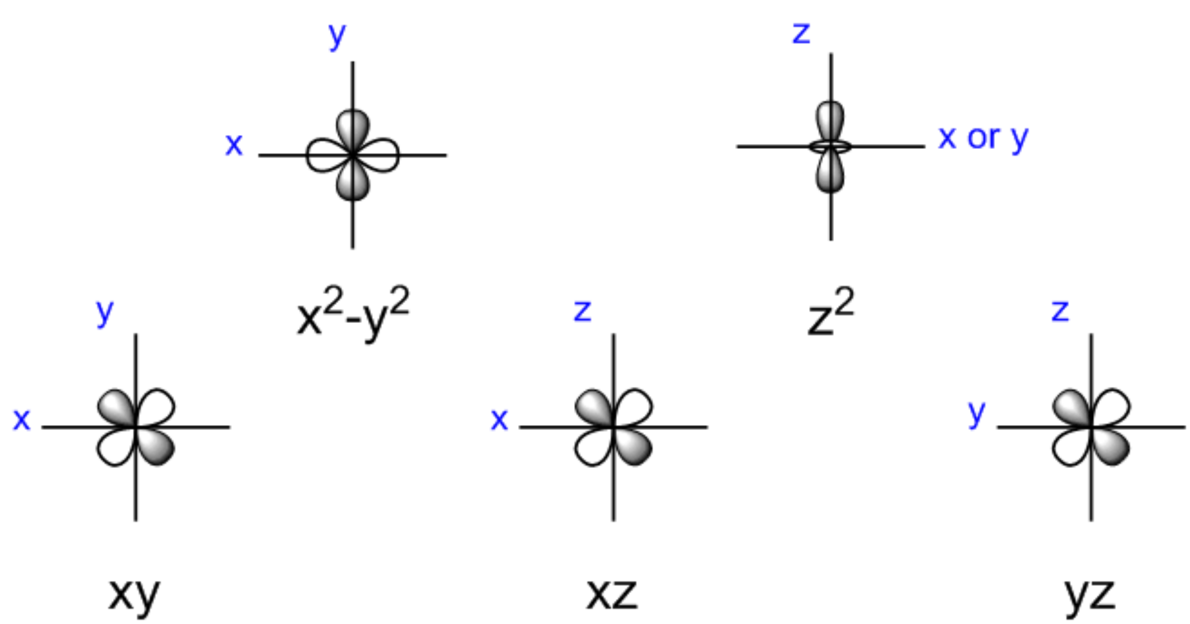
$$\boxed{Number\ of\ angular\ nodes = l}$$

$$\boxed{Number\ of\ radial\ nodes = n - 1 - l}$$

- There are two types of nodes: radial and angular.
- The number of angular nodes depends on the quantum number l.

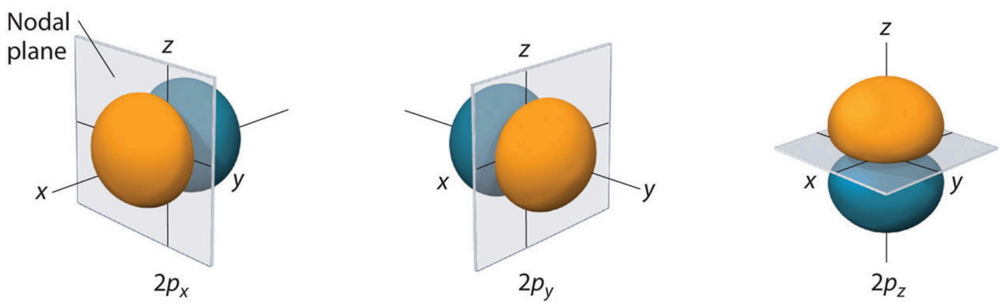
This time let's draw in the angular nodes for the 2p and 3d orbitals:

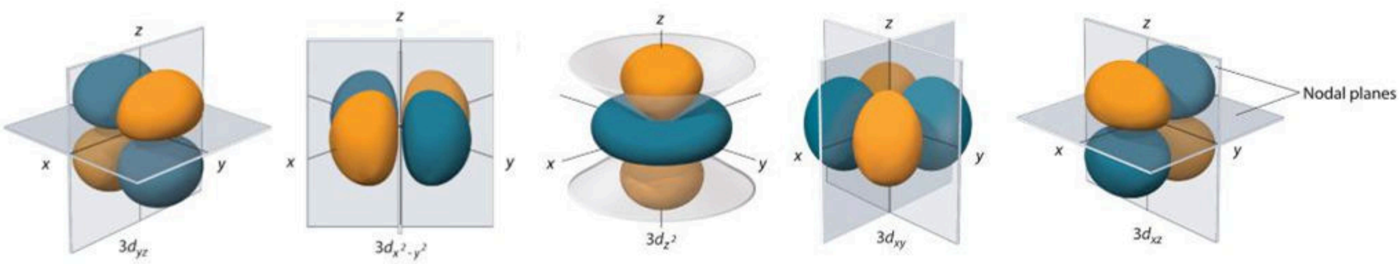




Another View of the Nodes

- **Nodes** are regions in space where the probability of finding an electron is zero.
- In p orbitals and d orbitals, we have **planar nodes** between the lobes of electron densities.
- We will take a look at s orbitals and their nodes soon...

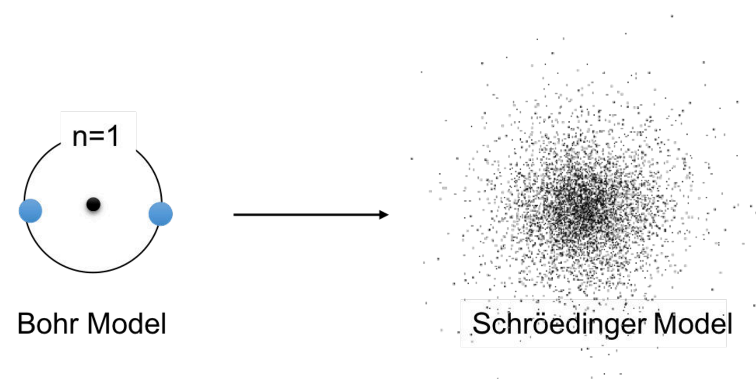




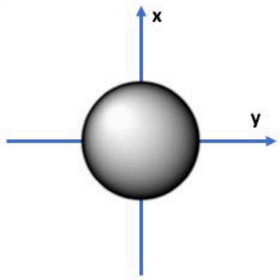
2.2.12

Wave Properties of Electrons

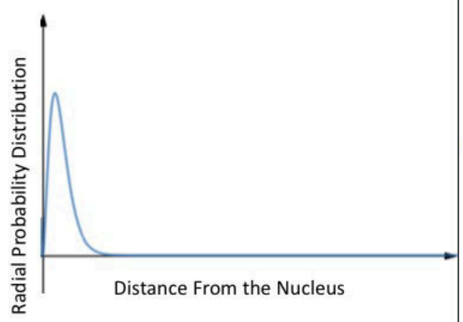
- Electrons have both particle and wave-like properties. Within an atom, electrons are standing waves, oscillating around the nucleus. As with all waves, there are nodes and phases.
- **Orbital:** mathematically derived region of probability in 3D space where an electron“may be” found



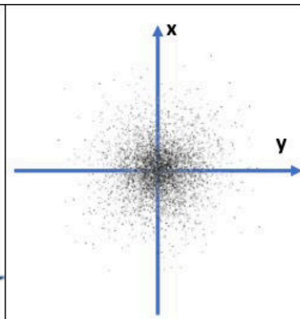
Representations of Atomic Orbitals



Boundary Diagram - What is the shape of the area in 3D space where the electron is found?



Radial Probability - What is the most probable distance the electron is from the nucleus?



Dot Density - Where can the electron be within the boundary?

Plotting the Radial Probability

The wavefunction is a mathematical function and it can be graphed like any other function.

WIZE TIP

We will not need to know any exact values of these graphs, just their **general shape.**

The only thing you will have need to “memorize” to do these problems is how to **draw each orbital.**

Although the wavefunction has no physical meaning, the square of the wavefunction is related to the probability of finding an electron at a given point in space.

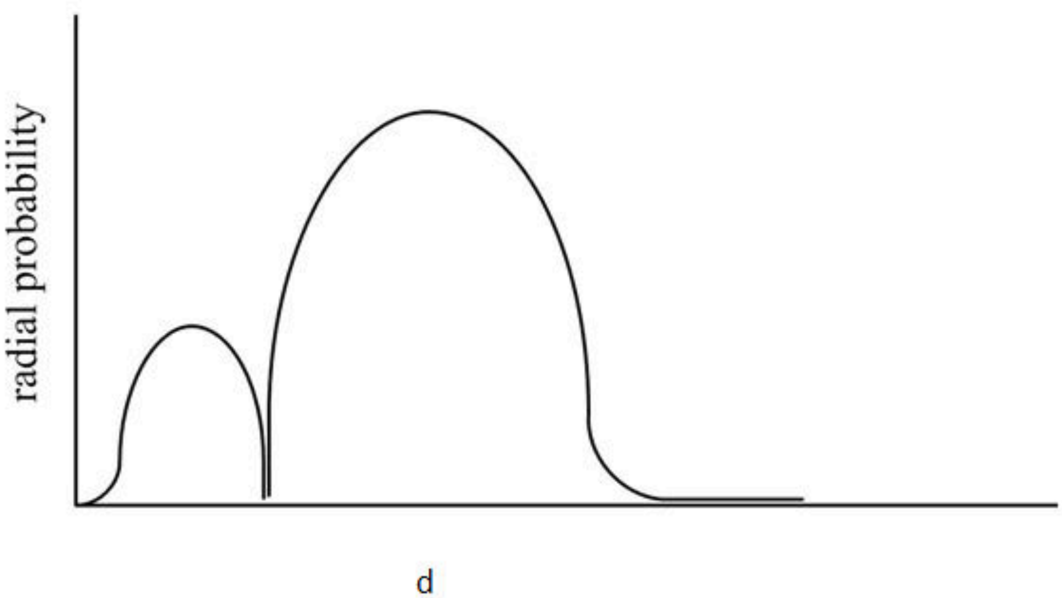
WIZE CONCEPT

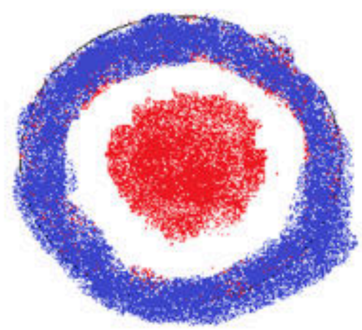
→ Each **radial node** will be indicated by a **zero point on a radial probability graph**

→ **Each graph begins at zero** (no probability of finding the electron in the nucleus)

→ **Each graph trails off to infinity** (no "edge" of an orbital")

Example: 2s Orbital





WATCH OUT!

The point where the **distance from the nucleus = 0 is NOT a radial node!**
This is a common mistake students make on exams.

Aside from that point, **every other point that touches the x axis represents a radial node!**

WIZE TIP

The most important piece of information we can get from the **radial probability plot** is the **number of radial nodes!**

2.2.14

Example: Calculate the Number of Radial and Angular Nodes

How many radial and angular nodes are present in the following orbitals?

a) 7p

b) 3d

c) 4s

2.2.15

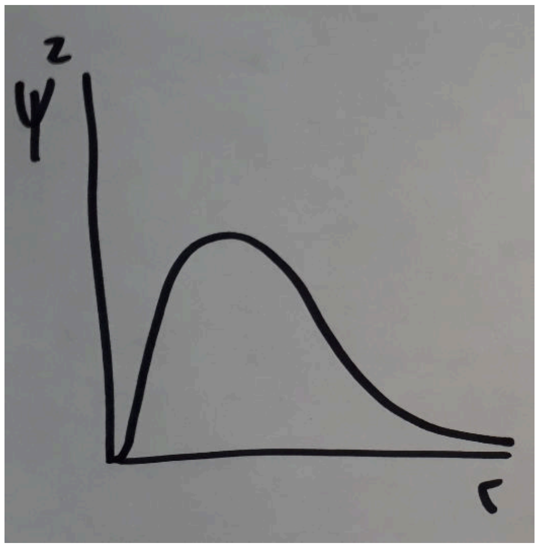
Example: Sketching Radial Probability

Sketch the radial probability of a $4d_{xy}$ orbital starting at r=0

2.2.16

Example: Identifying Orbitals from Probability Plots

Orbital X has 2 angular nodes and it's radial probability plot is shown below. Which of the following orbitals is most likely orbital X?



a) $3d_{z2}$

b)1s

c) $4p_x$

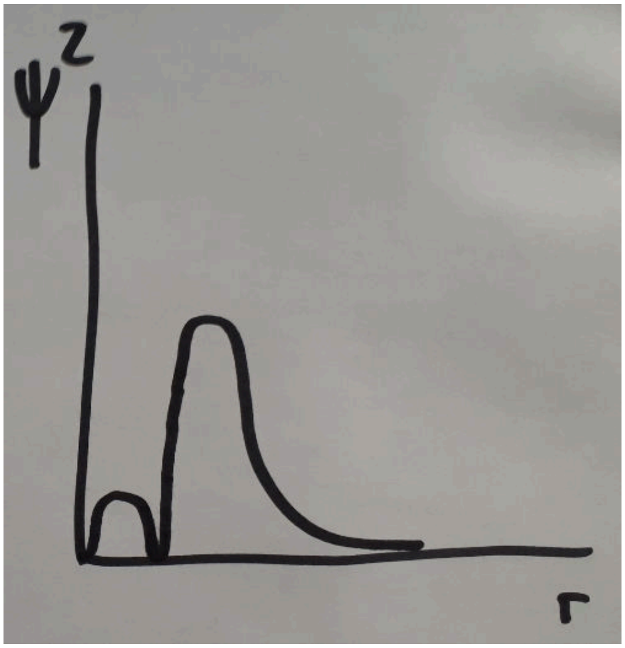
d) 3s

e) $2p_y$

2.2.17

Practice: Identifying Orbitals from Probability Plots

An orbital has only one angular node which lies along the yz plane. The radial probability plot is shown below, please identify the orbital.



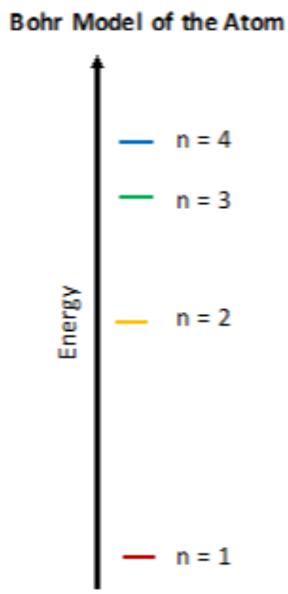
- $4d_{z2}$
- $4d_{xy}$
- $3p_z$
- 2s
- $3p_x$

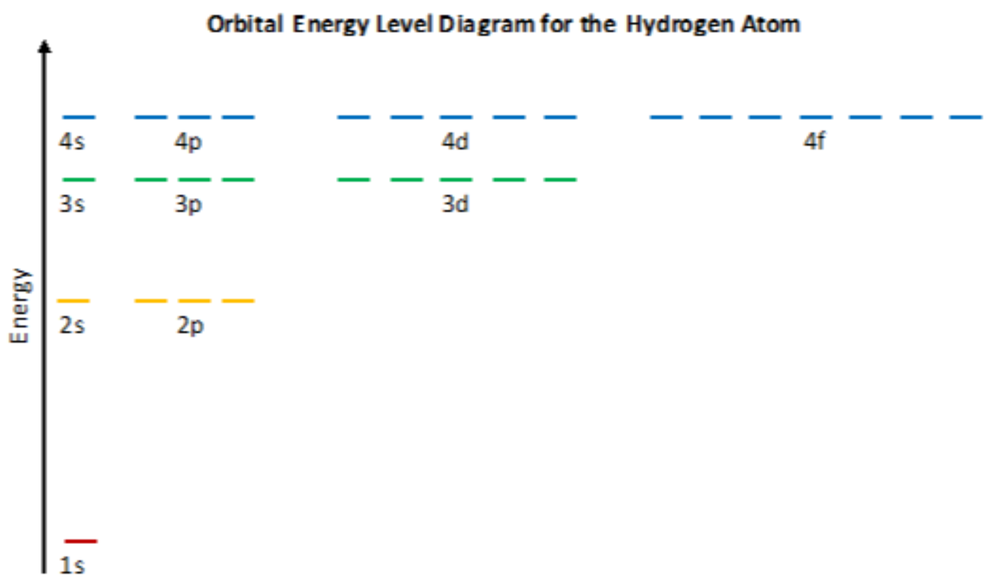
2.2.18

Relative Energies of Atomic Orbitals

One-Electron Species

- For a one-electron species like the hydrogen atom, there is no electron-electron repulsion. Therefore, all the subshells of the same n are **degenerate.**
- Orbitals at the **same energy level** are called **degenerate.**



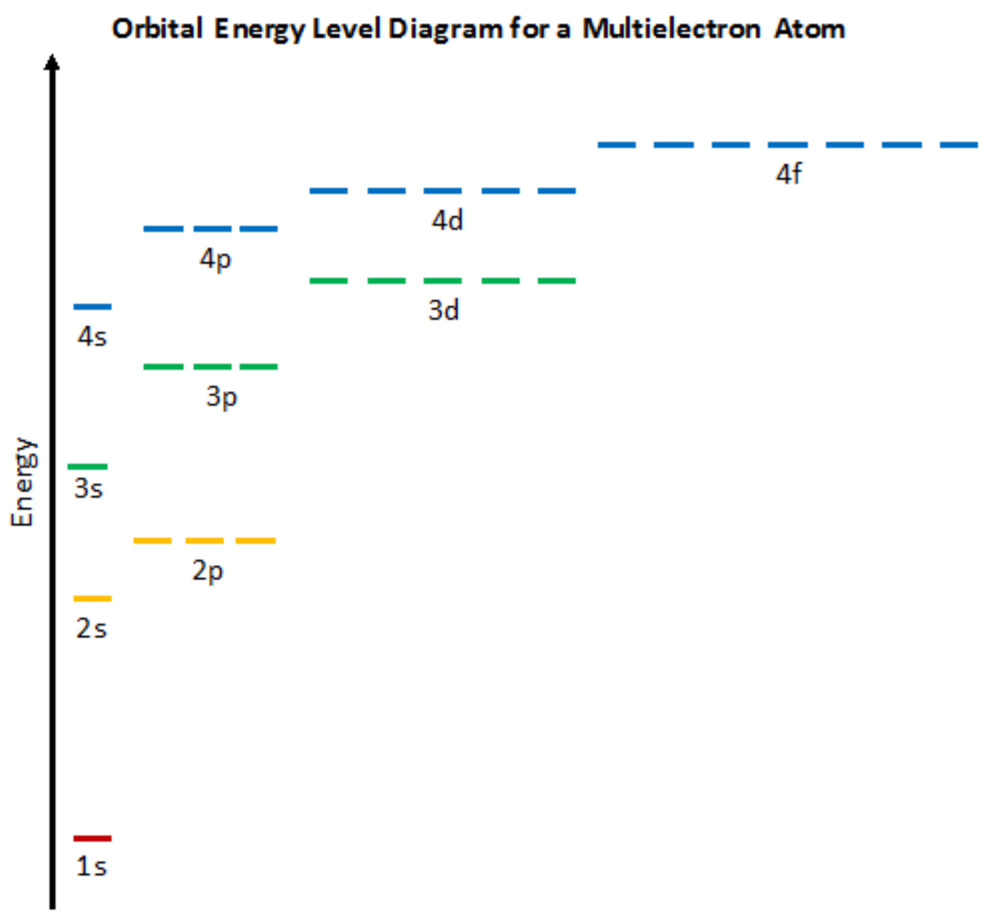


WIZE TIP

It will be helpful to understand the difference between a **shell**, a **subshell**, and an **orbital.**
We will label these in the diagram above!

Multi-Electron Species

- With multi-electron species, there are interactions between electrons (electron-electron repulsion)
- The energy level diagram of the orbitals looks like this:



WIZE CONCEPT

- S subshells hold a maximum of ______ electrons
- P subshells hold a maximum of ______ electrons
- D subshells hold a maximum of ______ electrons
- F subshells hold a maximum of ______ electrons

$$\boxed{\text{The total number of orbitals in a given shell is given by : } n^2}$$

Example:

If n=2, what is the total number of orbitals in that shell?

- ________, circle the _____ orbitals that have n=2 above

What if we wanted to know the maximum number of electrons we could have if n=2?

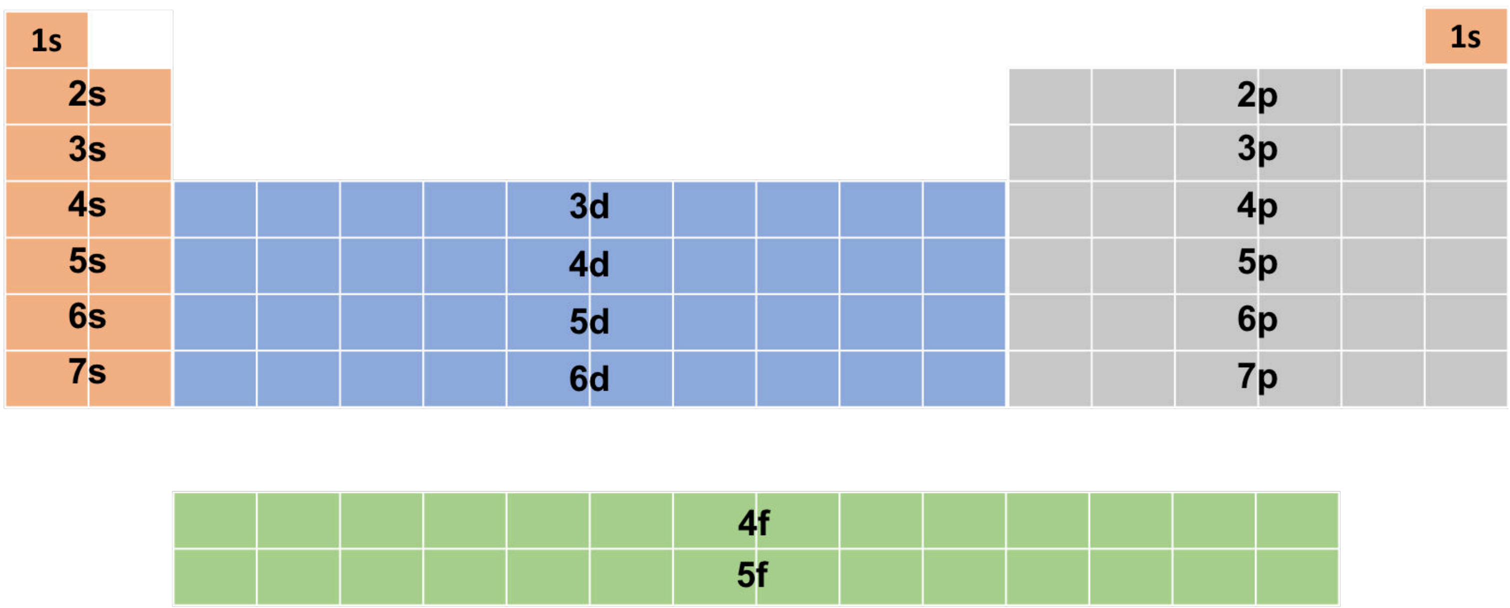
- We already found that total number of orbitals for this shell (_____)
- What is the maximum number of electrons we could have in each orbital? _____
- Therefore the maximum number of electrons possible when n=2 is _____

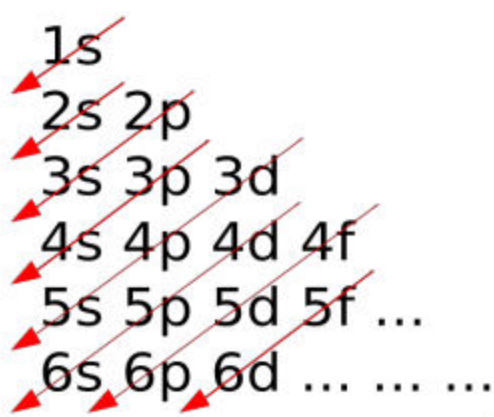
2.2.19

Rules for Orbital Filling

1) Aufbau Principle

Electrons will always occupy the **lowest available energy level first.**





WATCH OUT!

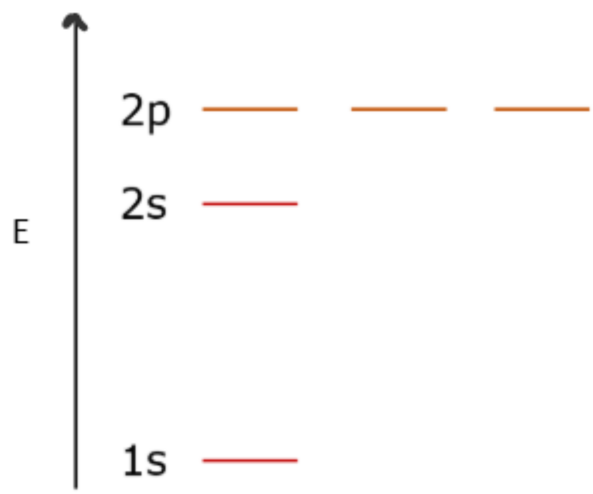
Memorize the order of orbital filling (or just familiarize yourself enough with the periodic table to know the order!): 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p...

2) Hund's Rule

Due to electron-electron repulsion, electrons will **fill orbitals of the same energy singly before pairing up**

Electrons don't want to be next to each other unless they have to be!

Example: Fill out the following orbital diagram for C



3) Pauli Exclusion Principle

No two electrons in an atom will have the same set of **4 quantum numbers.**

Quantum Number	Values	Interpretation
m_s (the spin quantum number)	-1/2 or +1/2	An electron behaves like a magnet that has one of two possible orientations, aligned either with the magnetic field or against it

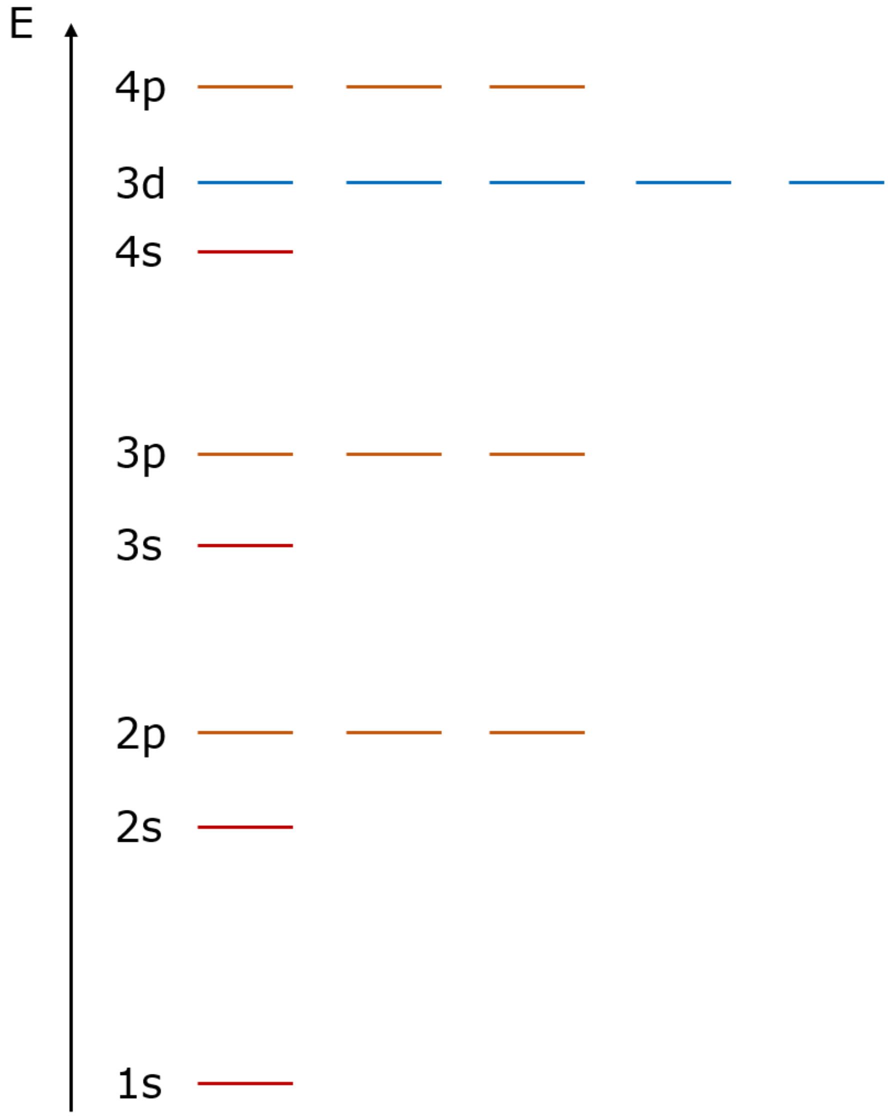
2.2.20

Example: Orbital Filling Diagrams

Draw the orbital diagram for oxygen.

Example: Orbital Filling Diagrams

Complete the atomic orbital diagram below for a neutral calcium atom.

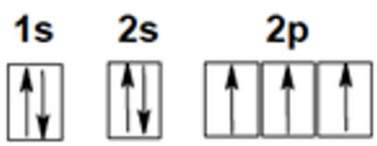


2.2.22

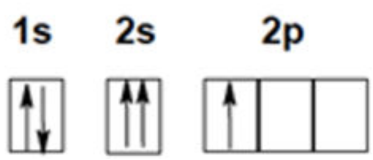
Practice: Orbital Filling Diagrams

Which of the orbital diagrams gives the correct electron configuration for an atom of boron, in the ground state?

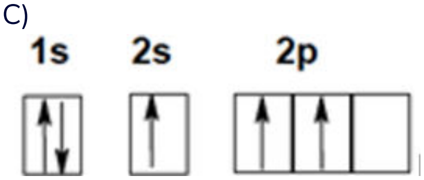
A)



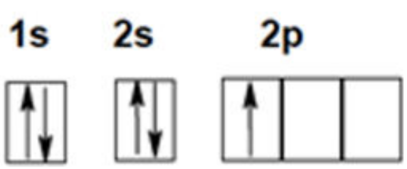
B)



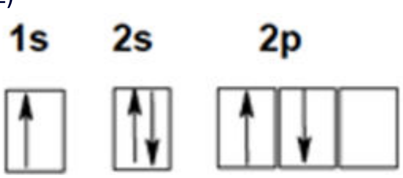
C)



D)



E)



2.2.23

Practice: Number of Orbitals

How many different atomic orbitals exist where n = 3?

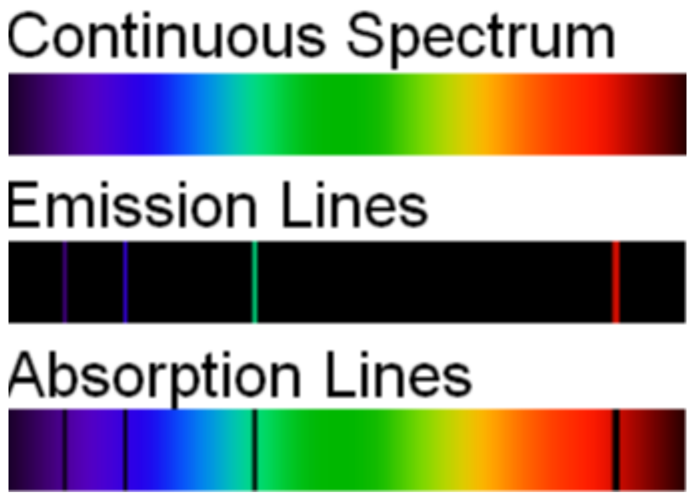
- 3
- 6
- 9
- 12, which
- 18

2.3 Energy Levels, Spectrum, and Ionization Energy of the Hydrogen Atom

2.3.1

The Basics of Absorption and Emission Spectra

Each atom has a **unique** absorption and emission spectrum.

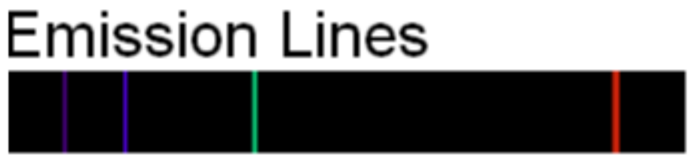


Absorption Spectra

- We get an absorption spectrum by transmitting electromagnetic radiation (light) through a substance.
- The **dark bands** in the spectra represent all of the **specific wavelengths of photons absorbed** by the atom's electrons

Emission Spectra

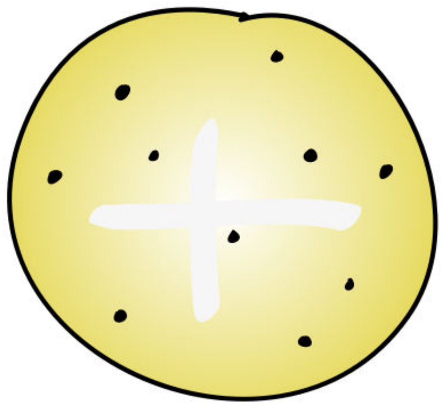
- We get an emission spectrum by measuring the electromagnetic radiation (light) that is emitted from a substance
- The background is black and the **coloured bands** represent the **wavelengths of photons that were emitted** from this atom's electrons



2.3.2

Bohr's Atomic Model

Before the Bohr model, scientists thought that electrons just assumed arbitrary energies and could be found anywhere in the atom.



If the pre-Bohr model was right, what would the emission spectra look like? Would they have lines or would they be continuous spectra? ____________________

Niels Bohr realized that the existing model at the time was inconsistent with emission spectra data... if electrons could be anywhere and have any energy, why did each element have a unique emission spectra with photons of light being emitted with specific energies and wavelengths?

Bohr realized that the **electrons must have discrete energies** in order to explain the emission spectra with photons with specific amounts of energy being released.

Bohr's Atomic Model Details

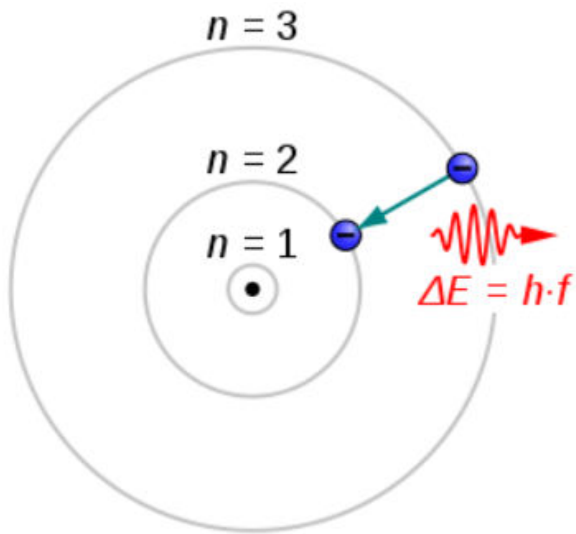


Photo by JabberWok / CC BY

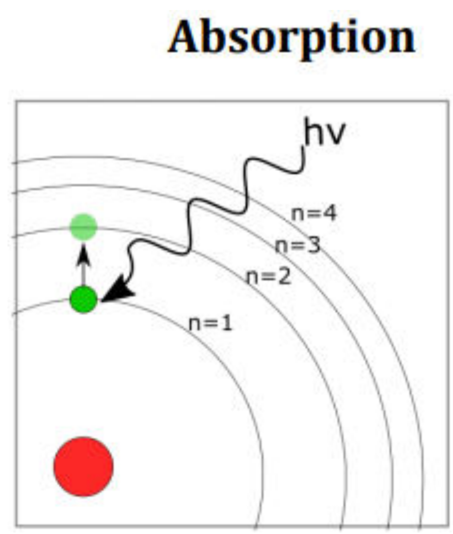
- Electrons orbit nucleus at **fixed distances** and in **circular paths** (like how planets orbit the sun!)
- Energy electrons have are **quantized** (only specific energies are allowed)
- Shell's distance from nucleus determines energy (n=1, 2, 3, etc, **higher n=higher energy shell**)
- **Gap between levels shrinks as n grows**
- Bohr model is only accurate for atoms with **1 electron** ("**Bohr atoms**")
  - The Bohr model explains why only photons with certain energies are observed in the emission spectra!

Analogy: Staircase vs Ramp

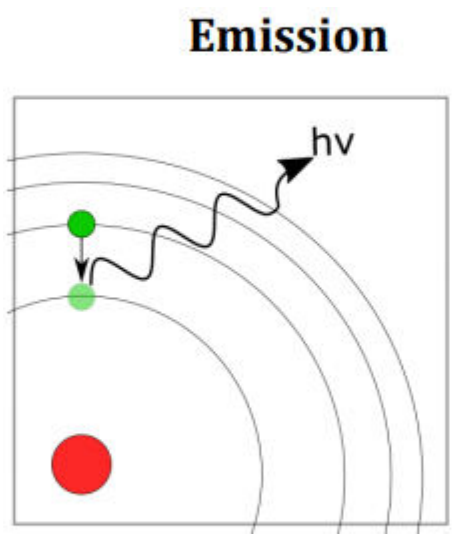
- **Ramp->**if you're walking up a ramp, you could be 12cm, 17cm, or 20cm above the ground. (Pre-Bohr Model)
- **Staircase->** if each step is 17cm high, you can never be standing 12 cm or 20cm above the ground (Bohr Model)
    - The **stairs are discrete units** just like **shells in an atom are discrete units.**
    - And just like how we can only be **certain heights above the ground when we take the stairs**, **electrons can only be in certain shells with certain energies (can't be between two shells)**

2.3.3

Absorption vs Emission of a Photon



- Atom **absorbs photon** with **energy** (E=hv) **equal to gap between levels**
- Electron is promoted to **higher energy level** (ex: n = 1 → n = 2)
- **Greater Δn = higher energy**



- Atom **emits photon** with **energy equal to gap between levels**
- Electron relaxes to **lower energy level** (ex: n = 4 → n = 2)

Example: Understanding Absorption Spectra and the Bohr Model

To collect an absorption spectra of a gaseous mixture a gas is irradiated with light from each wavelength in a given range and the amount of light which passes through the sample is recorded. Some of the light at certain wavelengths is absorbed by the sample and these wavelengths become peaks in the absorption spectra.

1) An absorption event occurs when...

a) The atom is destroyed

b) The nuclear charge increases by 1

c) An isotope ejects an electron

d) An electron is excited from one energy level to a higher energy level

e) None of the above

2) The wavelength of light which is absorbed corresponds with ...

a) The difference in energy between the energy levels which the electrons occupy

b) The energy level of the orbit the electron was initially in

c) The energy level of the orbit the electron ends up in

d) The distance between the two orbits the electron moves between

e) None of the above

2.3.5

Example: Electron Transitions

An electron in the n = 4 shell of an excited hydrogen atom relaxes to it's ground state. Which of the following is true?

a) A photon is absorbed and with energy equal to the energy of the ground state

b) A photon is absorbed with energy slightly less than the difference in energy between n = 4 and n = 1

c) A photon is absorbed with energy equal to the difference in energy between n = 4 and n = 1

d) A photon is released with energy equal to the energy of the ground state

e) A photon is released with energy slightly less than the difference in energy between n = 4 and n = 1

f) A photon is released with energy equal to the difference in energy between n = 4 and n = 1

2.3.6 **Bohr Model Equations**

Bohr Model Equations

There are a few important equations we should be familiar with that are related to the Bohr model. **These equations can only be used for one electron systems!** Recall that **Bohr's model** was only able to explain species with **1 electron.**

Calculating the Energy of a Specific *n* Level

$$\boxed{E = \frac{\left(-2.178 \times 10^{-18}\mathrm{J}\right)\left(Z^2\right)}{n^2}}$$

*where **n** is the shell # and it can be 1, 2, 3...etc*
***E** represents the **energy** of a specific n level*
***Z=atomic # (# of protons)** in an atom (ex. Hydrogen has 1 proton, Z=1, He+ has 2 protons, Z=2)*

Calculating the Energy Difference Between n Levels

$$\boxed{\Delta E = -2.178\ \times 10^{-18} J\left(Z^2\right)\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)}$$

*where **ΔE** represents the **difference in energy** between 2 energy levels*
***Z=atomic # (# of protons)** in an atom*
*n_f is referring to the **final energy level***
*n_i is referring to the **initial energy level***

- If **ΔE is positive**, energy was (absorbed/released) ____________ and the electron should go to a (higher/lower) _________ energy level
- If **ΔE is negative,** energy was (absorbed/released) ____________ and the electron should go to a (higher/lower) _________ energy level

WATCH OUT!

Be very careful when entering n_f and n_i values into this equation!

Example: If we are told energy is **absorbed** and an electron goes from **n=1 to n=2**

- **n=1 is the initial energy level, n=2 is the final energy level.**

Example: If we are told energy is **released** as an electron goes from **n=3 to n=2**

- **n=3 is the initial energy level, n=2 is the final energy level.**

Calculating the Wavelength of the Photon Absorbed or Emitted

Method 1:

- Use the above equation to solve for ΔE.
- Then use $E = \frac{hc}{\lambda}$ to solve for the wavelength of the photon
  - Even if ΔE is negative, use the positive value to solve for the wavelength so you **get a positive answer for wavelength** (negative wavelength doesn't make sense)

Method 2:

- Use the following equation and then take the inverse to solve for λ

$$\boxed{\frac{1}{\lambda} = R_H\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)}$$

λ=wavelength *of the photon (**m**) -always **enter answer as a positive value!***

R_H=Rydberg constant=1.097x$10^7 m^{-1}$

n_i=initial energy level, n_f=final energy level

2.3.7

Example: Releasing or Absorbing Energy and How That Differs For Different Atoms

An electron in a one-electron species transitions from the n = 4 state to the n = 2 state.

a) Is energy released or absorbed?

b) Would more energy be absorbed/released from a sample of Hydrogen atoms or Li^{2+} ions?

Example: Excitation and Relaxation in Atoms

Calculate the wavelength, in nm, of a photon released when an electron in a hydrogen atom relaxes from n = 3 to n = 1. What is the energy of the photon if the electron is re-absorbed from n = 1 to n = 3?

2.3.9

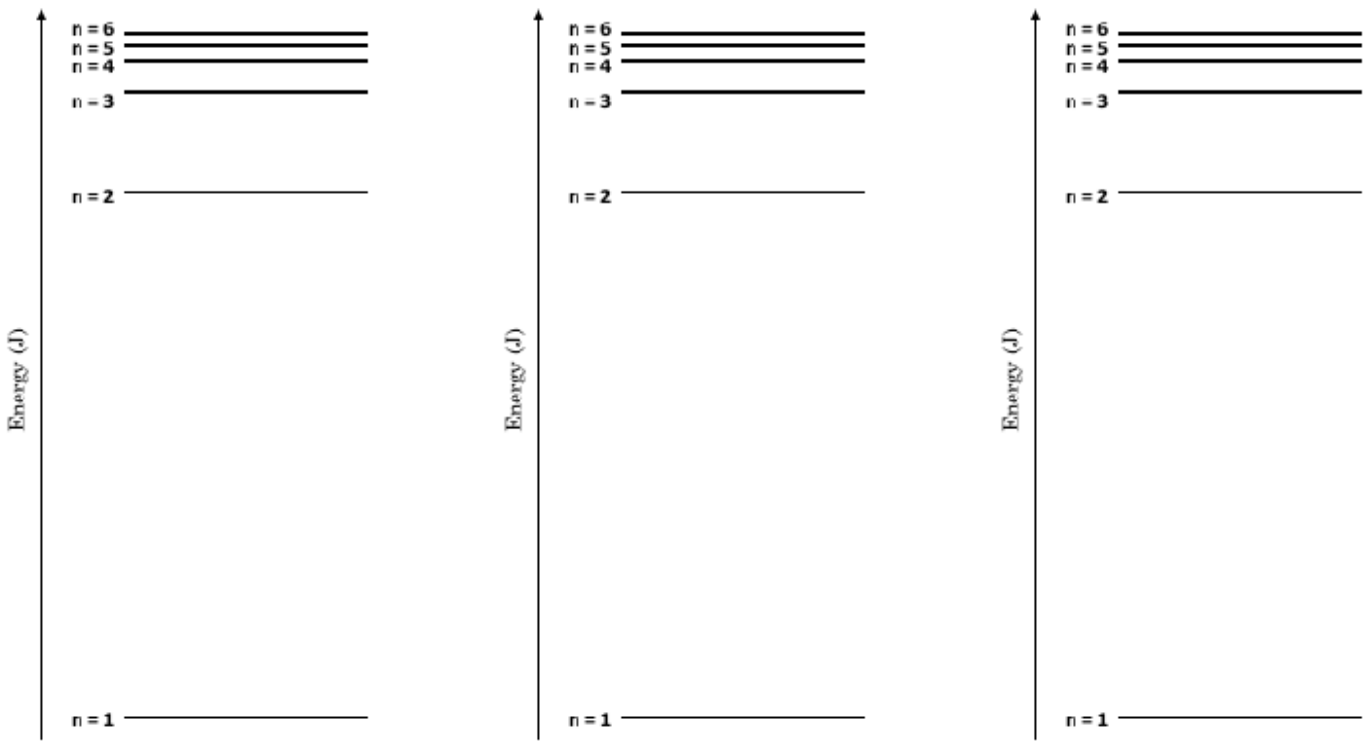
Practice: Using Bohr Model Equations to Determine an Energy Level

A photon of light was released by a hydrogen atom with an energy of 1.94 x 10^{-18} J. If this transition ended at n=1, from which level did this transition begin?

- n=1
- n=2
- n=3
- n=5
- n=6

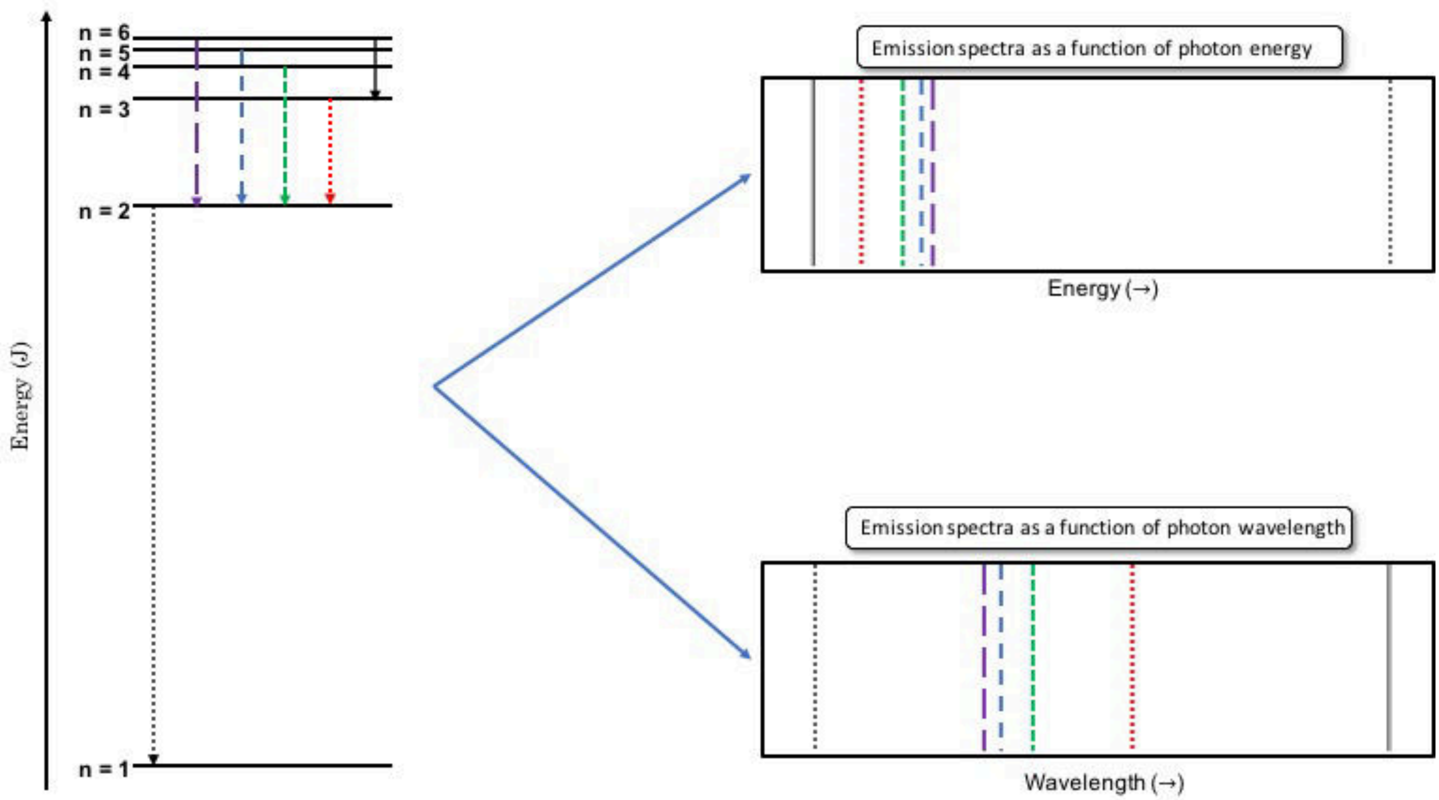
Atomic Line Spectra

Emission and Absorption



Note: **n=1 is ground state**, all other n levels are excited states

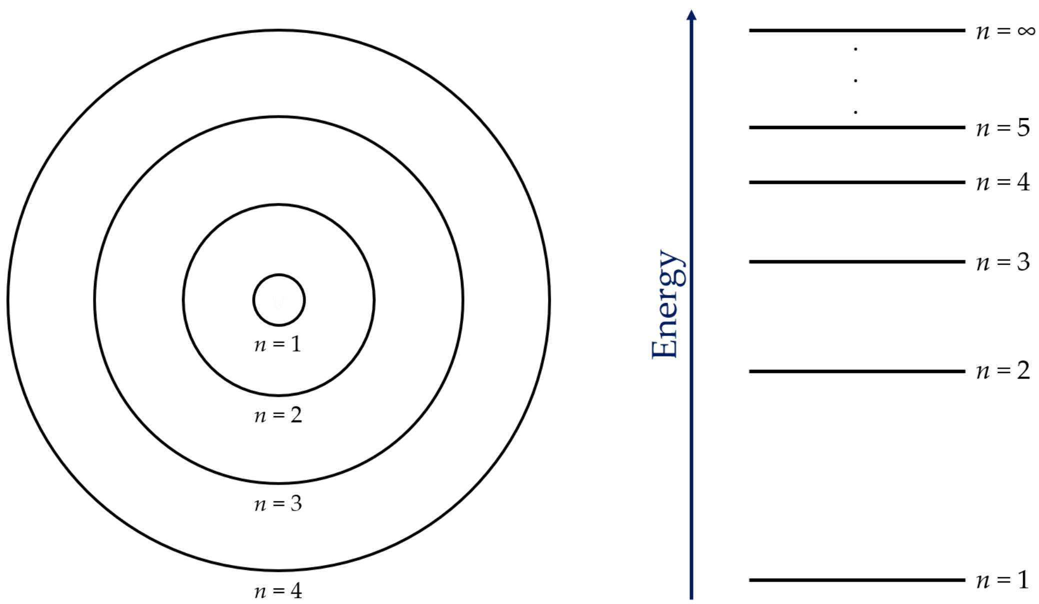
Emission Spectra

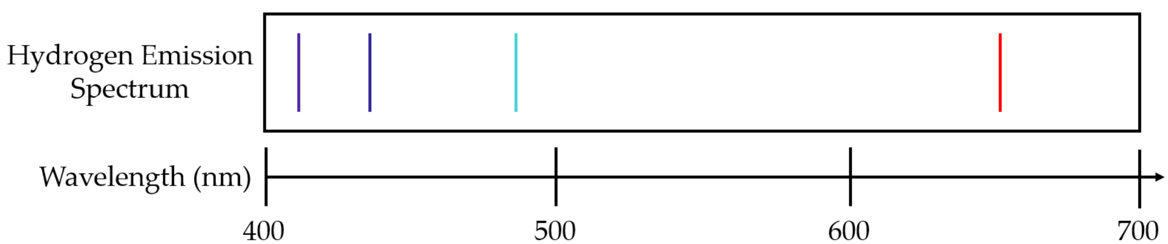


2.3.11

Relating Electronic Transitions to Emission Spectra

Bohr Model of the Hydrogen Atom





Hydrogen Emission Spectral Series:

Hydrogen Emission Spectrum Series Name	n_f
Lyman	1
Balmer	2
Paschen	3
Brackett	4
Pfund	5

Example: Removing an Electron

What wavelength of light corresponds to the 4^{th} ionization energy of Beryllium?

Can start by solving for the energy difference for this transition:

Now you can solve for the wavelength that the energy difference corresponds to:

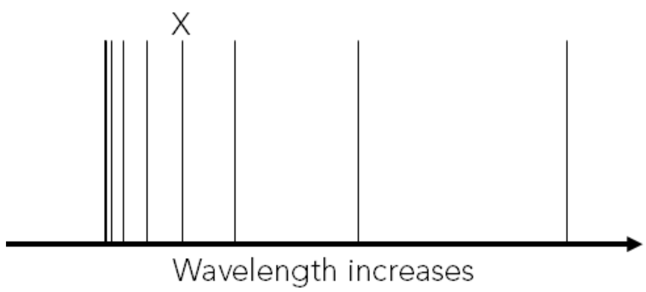
Practice: Understanding Emissions

Which of the following transitions would emit the shortest wavelength photon?

- n=2 to n=1
- n=1 to n=6
- n=5 to n=1
- n=7 to n=2
- n=1 to n=4

2.3.14

You shine light through a sample of gaseous atomized hydrogen and take an absorption spectrum as shown below. Only transitions to or from the ground state (n=1) appear in the spectrum. Which of these energy-level transitions correspond to the line labelled X in the absorption spectrum?



- n = 1 to n = 2
- n = 2 to n = 1
- n = 1 to n = 4
- n = 1 to n = 5
- n = 4 to n = 1
- n = 5 to n = 1

2.4 Quantum Theory of the Hydrogen Atom

2.4.1

Bohr Model Vs Quantum Model

What was good about the Bohr model and theory?

✔️ Could predict line spectra for hydrogen and other ions with **1 electron** (ex. Li^{2+})

❌ But it couldn't predict line spectra for species with more than 1 electron

Couldn't explain how atoms were bound together

Didn't explain why only certain orbits were allowed

❌ Tries to be too precise to describe where electrons are found

Bohr Theory was replaced by Quantum Mechanics!

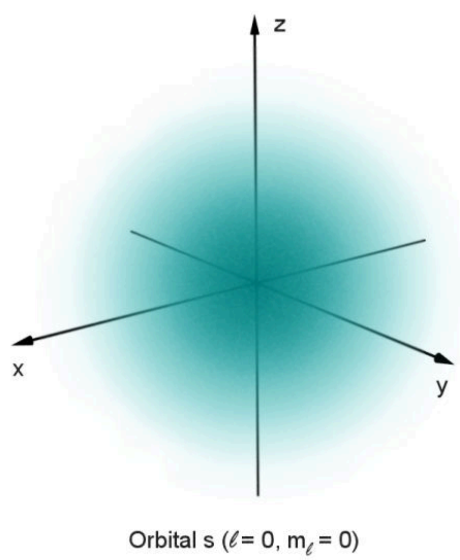
✔️ Quantum model can predict atomic spectra for **many-electron atoms**

Quantum mechanics provides us with **probabilities where electrons are most likely to be found**

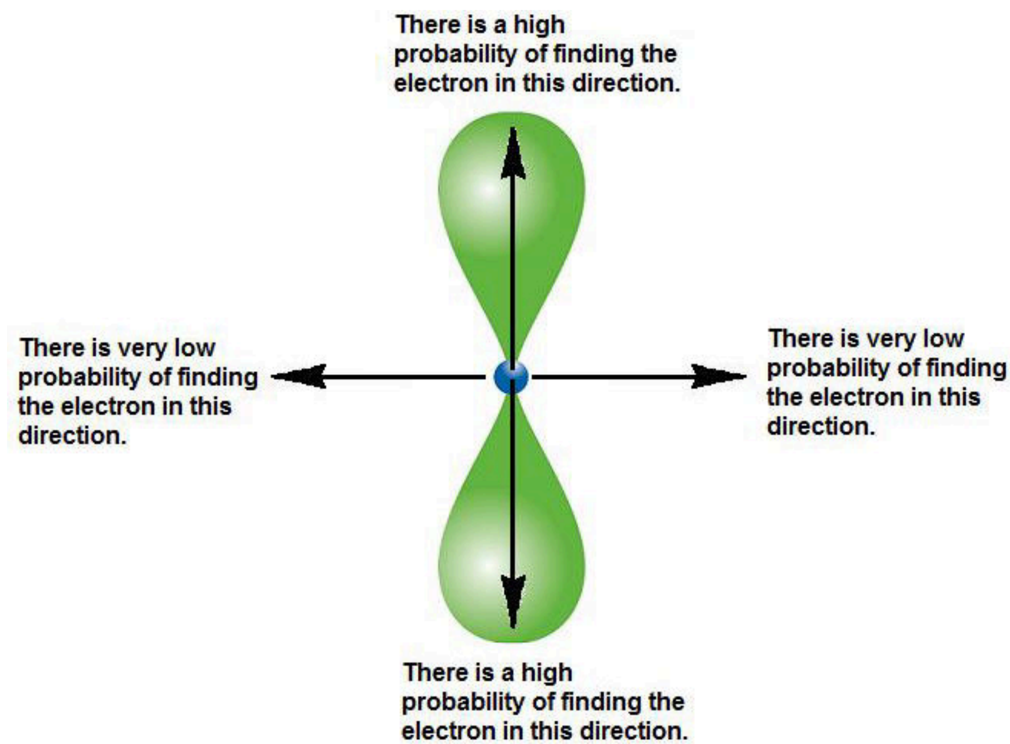
- Provides a 3D region where electrons are most likely to be found (instead of circular paths in Bohr model)

✔️ Adds on to the Bohr model:

  - Electrons are still only found at discrete energy levels with specific amounts of energy (quantized)
  - In this model, **each electron has 4 quantum #s that are unique to that electron** and describe it.



Electron Density Diagram


There is a high probability of finding the electron in this direction.
There is very low probability of finding the electron in this direction.
There is very low probability of finding the electron in this direction.
There is a high probability of finding the electron in this direction.

2.5 Electron Spin: A Fourth Quantum Number

2.6 Electron Configurations

2.6.1

Electron Configurations of Atoms

Group → ↓Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba		72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra		104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Ms	116 Lv	117 Ts	118 Og

Lanthanides	57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
Actinides	89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

- lable blocks
- describe the n number for each block

Simple Electron Configurations

Li:

C:

Ne:

Group → ↓Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba		72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra		104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Ms	116 Lv	117 Ts	118 Og

Lanthanides	57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
Actinides	89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

Shorthand Notations

Write the name of the **previous noble gas and then fill in the rest of the electron configuration.**

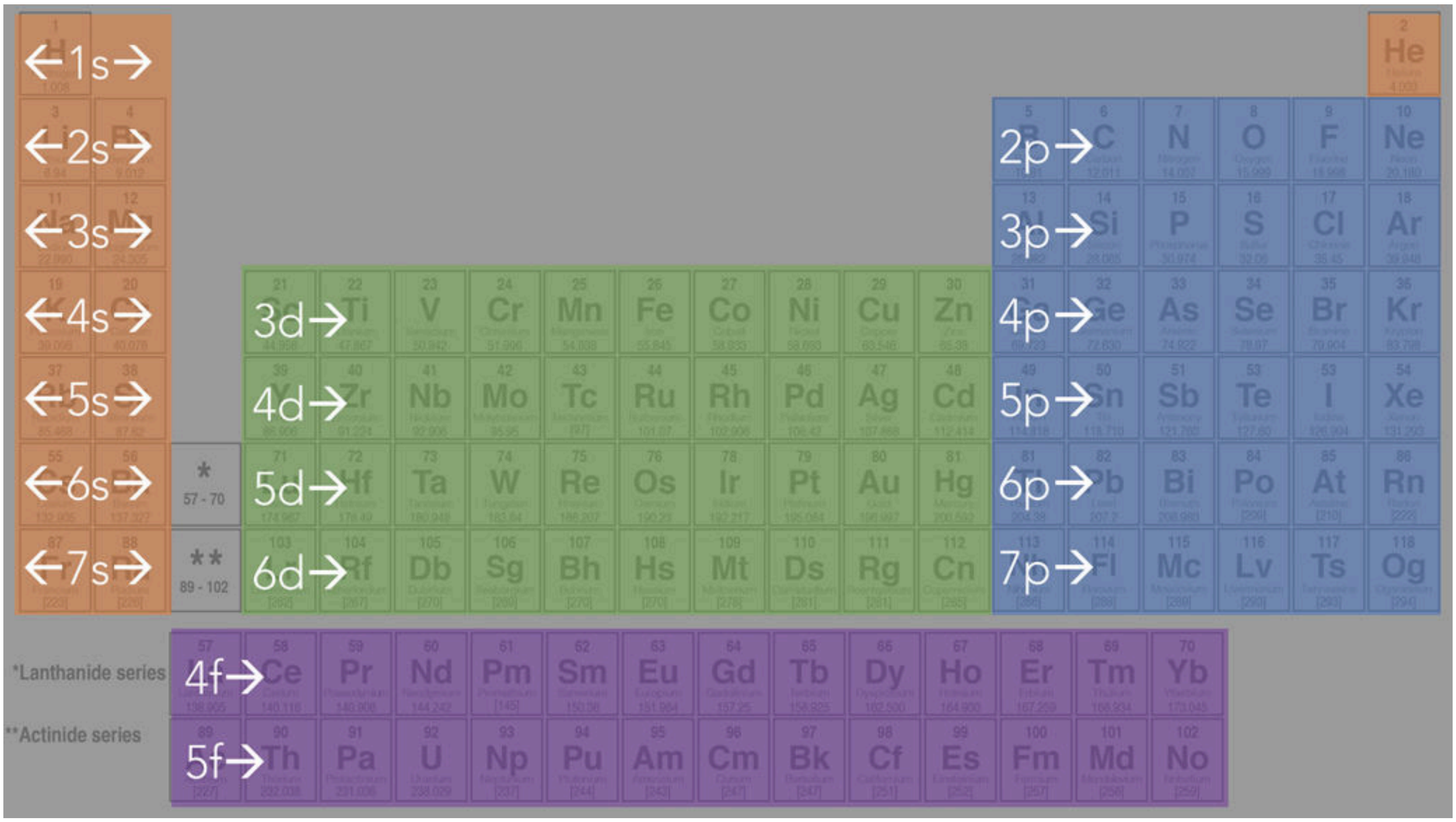
O:

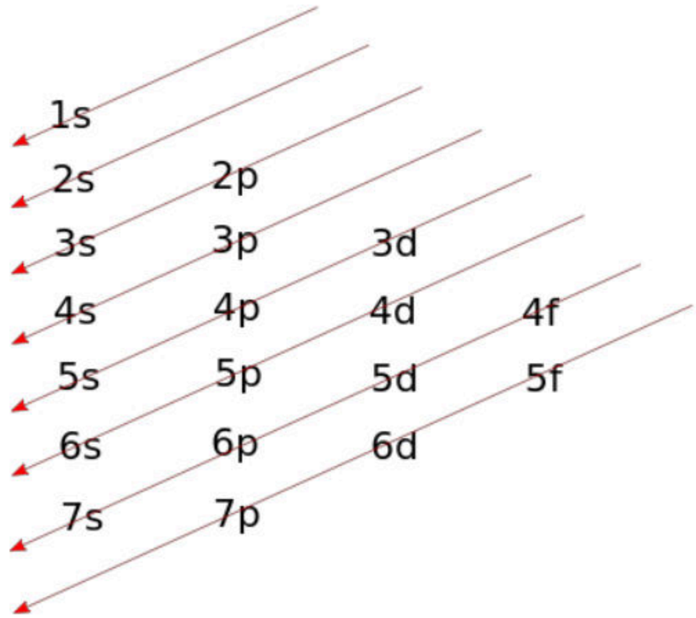
Fe:

Ge:

2.6.2

Electron Configurations Cheatsheet




1s
2s 2p
3s 3p 3d
4s 4p 4d 4f
5s 5p 5d 5f
6s 6p 6d
7s 7p

2.6.3

Electron Configuration Exceptions

Transition metals have their valence electrons in the d subshell.
It is a lot **more favourable** (and stable) to have **d shell either half full or totally full instead of just partially full.**

	Cr:	Cu:
Expected:	4s 3d [Ar] $4s^2 3d^4$	4s 3d [Ar] $4s^2 3d^9$
Seen:	4s 3d [Ar] $4s^1 3d^5$	4s 3d [Ar] $4s^1 3d^{10}$

WIZE TIP

The other exceptions that follow this same trend are: Mo, Ag, and Au! (Note they are in the same groups that Cr and Cu are in!)

2.6.4

Example: Writing Out Electron Configurations

Part 1:

Write the full electronic configurations for the following elements:

P:

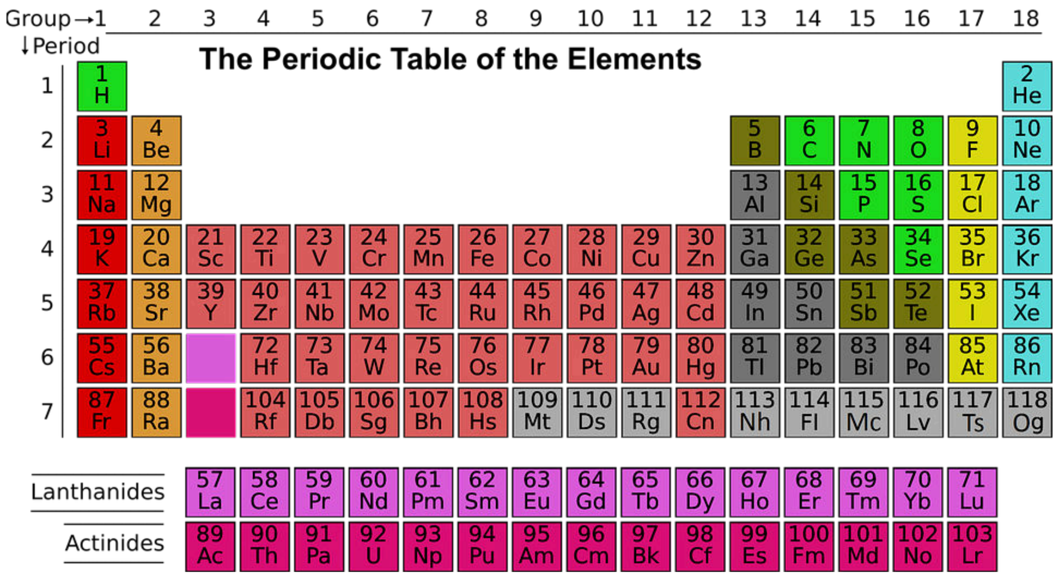
Se:

Part 2:

Write the short electron configurations for the following elements:

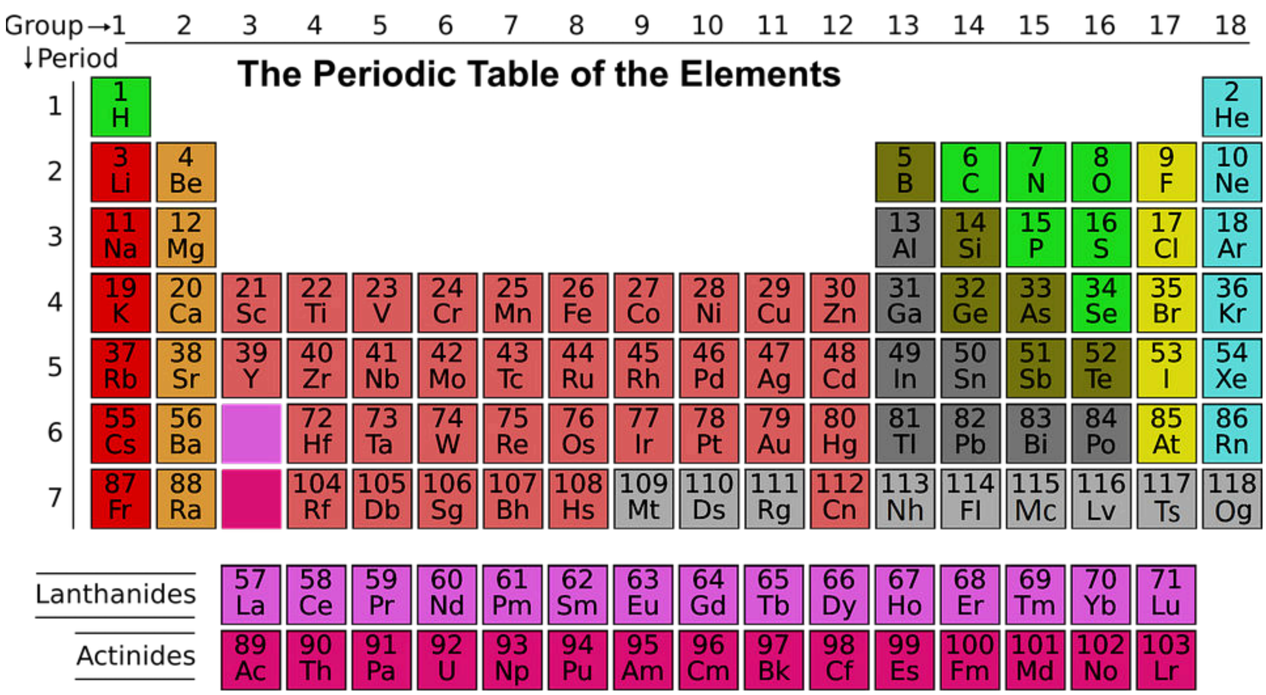
Zn:

Zr:

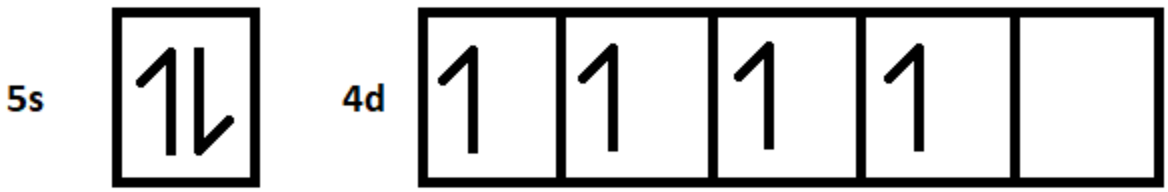

Group→1 2 3 4 5 6 7 8 9 10 11 12 13 14 15 16 17 18
↓Period
The Periodic Table of the Elements
1 1 H 2 He
2 3 Li 4 Be 5 B 6 C 7 N 8 O 9 F 10 Ne
3 11 Na 12 Mg 13 Al 14 Si 15 P 16 S 17 Cl 18 Ar
4 19 K 20 Ca 21 Sc 22 Ti 23 V 24 Cr 25 Mn 26 Fe 27 Co 28 Ni 29 Cu 30 Zn 31 Ga 32 Ge 33 As 34 Se 35 Br 36 Kr
5 37 Rb 38 Sr 39 Y 40 Zr 41 Nb 42 Mo 43 Tc 44 Ru 45 Rh 46 Pd 47 Ag 48 Cd 49 In 50 Sn 51 Sb 52 Te 53 I 54 Xe
6 55 Cs 56 Ba 72 Hf 73 Ta 74 W 75 Re 76 Os 77 Ir 78 Pt 79 Au 80 Hg 81 Tl 82 Pb 83 Bi 84 Po 85 At 86 Rn
7 87 Fr 88 Ra 104 Rf 105 Db 106 Sg 107 Bh 108 Hs 109 Mt 110 Ds 111 Rg 112 Cn 113 Nh 114 Fl 115 Mc 116 Lv 117 Ts 118 Og
Lanthanides 57 La 58 Ce 59 Pr 60 Nd 61 Pm 62 Sm 63 Eu 64 Gd 65 Tb 66 Dy 67 Ho 68 Er 69 Tm 70 Yb 71 Lu
Actinides 89 Ac 90 Th 91 Pa 92 U 93 Np 94 Pu 95 Am 96 Cm 97 Bk 98 Cf 99 Es 100 Fm 101 Md 102 No 103 Lr

Example: Condensed Electron Configuration

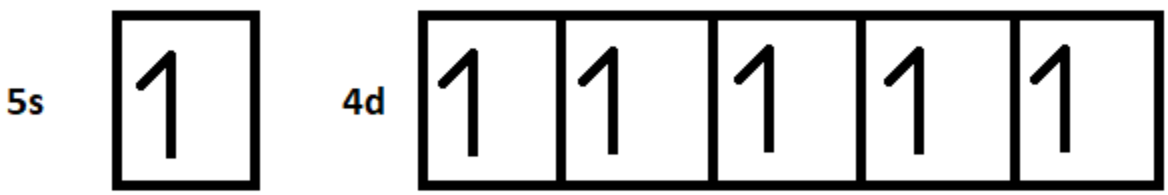
What is the condensed electronic configuration for Mo?



We ignore everything before the previous noble gas (Kr), which means we are only working with the **5s** and **4d** blocks. Let's fill them up:



But there is something special about Mo. It is in group 6. This group tends to **fill (or in this case half-fill) the d-block** at the expense of the s block.



Therefore, the electronic configuration is **[Kr]$5s^1 4d^5$.**

2.6.6

Practice: Identify the Neutral Element

What neutral element is represented in the example below?

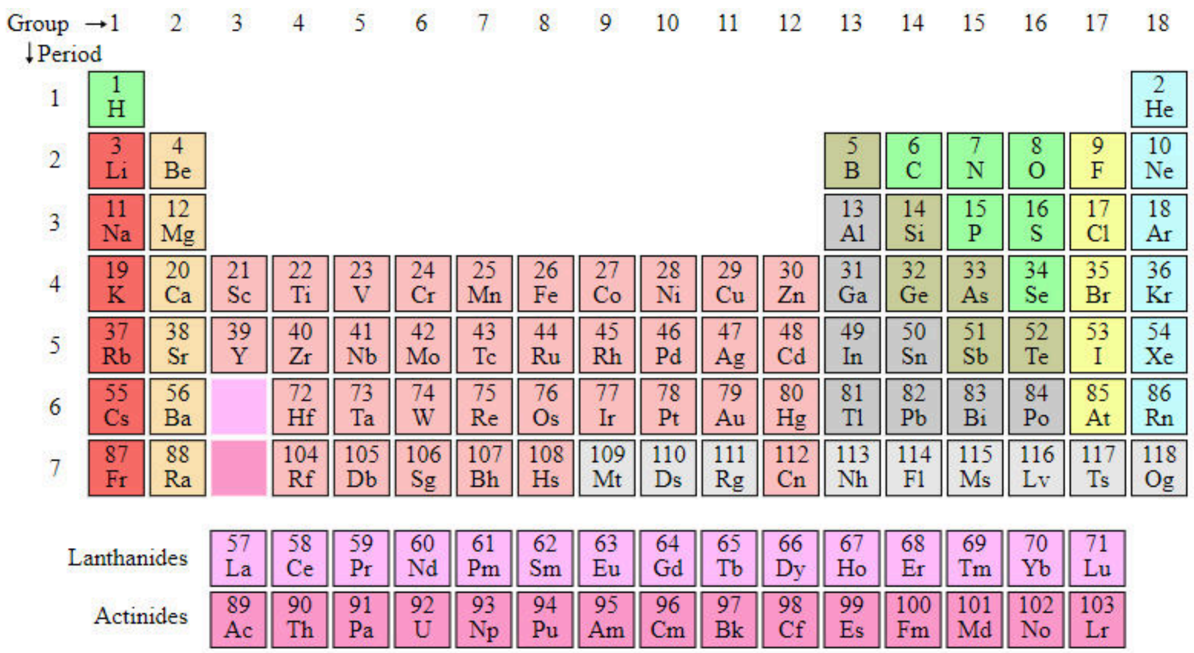
$1s^2 2s^2 2p^6 3s^2 3p^5$ became $[Ne]3s^2 3p^5$

- Al
- Ne
- P
- Cl

2.6.7

Practice: Electron Configuration

The ground state electronic configuration for At is



A) $[Xe]6s^24f^{14}5d^{10}6p^5$

B) $[Xe]6s^25d^{10}6p^5$

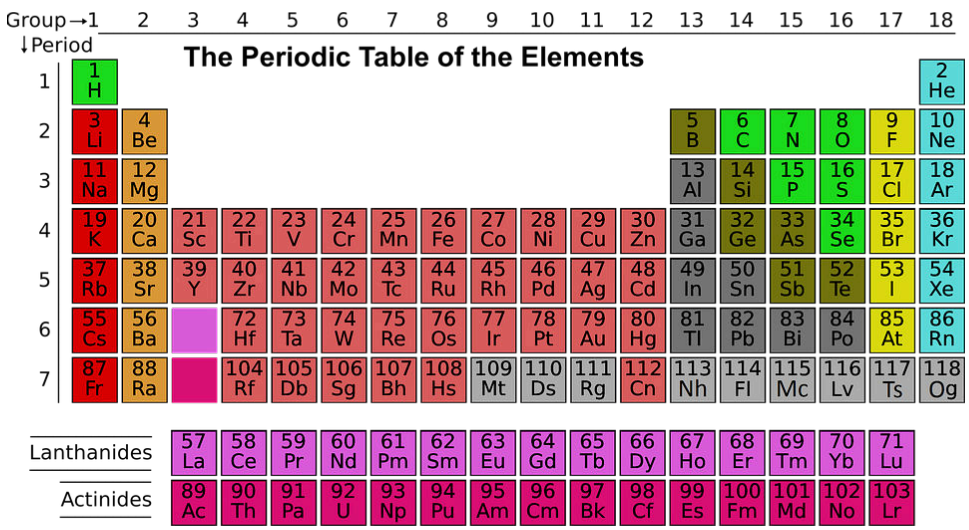
C) $[Xe]6s^26p^5$

D) $[Xe]6s^24f^{14}5d^{10}4p^3$

2.6.8 Practice

Practice: Impossible Electron Configuration

Which of the following electronic configurations is not possible?



- [Ar]$4s^2 4d^6$
- $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$
- [Kr]$5s^1$
- $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^8$

2.6.9

Electron Configuration of Ions of Main Group Elements

We will now look at how to write the electron configurations for ions!

Writing the electron configurations for anions (**negative ions**) are simple.

1. Write out the electron configuration for the **neutral element**
2. Fill in the desired number of electrons by **adding electron(s) to the highest energy subshell**

Writing the electron configurations for cations (**positive ions**) usually confuses students more, but there is a simple thing to remember to get these questions right every time!

1. Write out the electron configuration for the **neutral element**
2. Remove the desired number of electrons by **first removing electrons from the highest n level and highest energy subshell!**

WIZE CONCEPT

The higher the n level, the higher the energy.
Example: **2**s has a higher energy than **1**s

In order of **increasing energy for the subshells**, we write:
s < p < d < f
Example: 2**p** has a higher energy than 2**s**

F: $1s^22s^22p^5$ **[He] $2s^22p^5$**	2s	2p
F^-: $1s^22s^22p^6$ **[Ne]**	2s	2p

What is the electron configuration for F^+?

Solution available online

WIZE TIP

The most stable ions of atoms are **isoelectronic** with the noble gases or have filled shells. **Isoelectronic species** have the same number of electrons and the same electron configuration. The term "isoelectronic" is commonly seen on exams.

Example: F^- is isoelectronic with Ne!

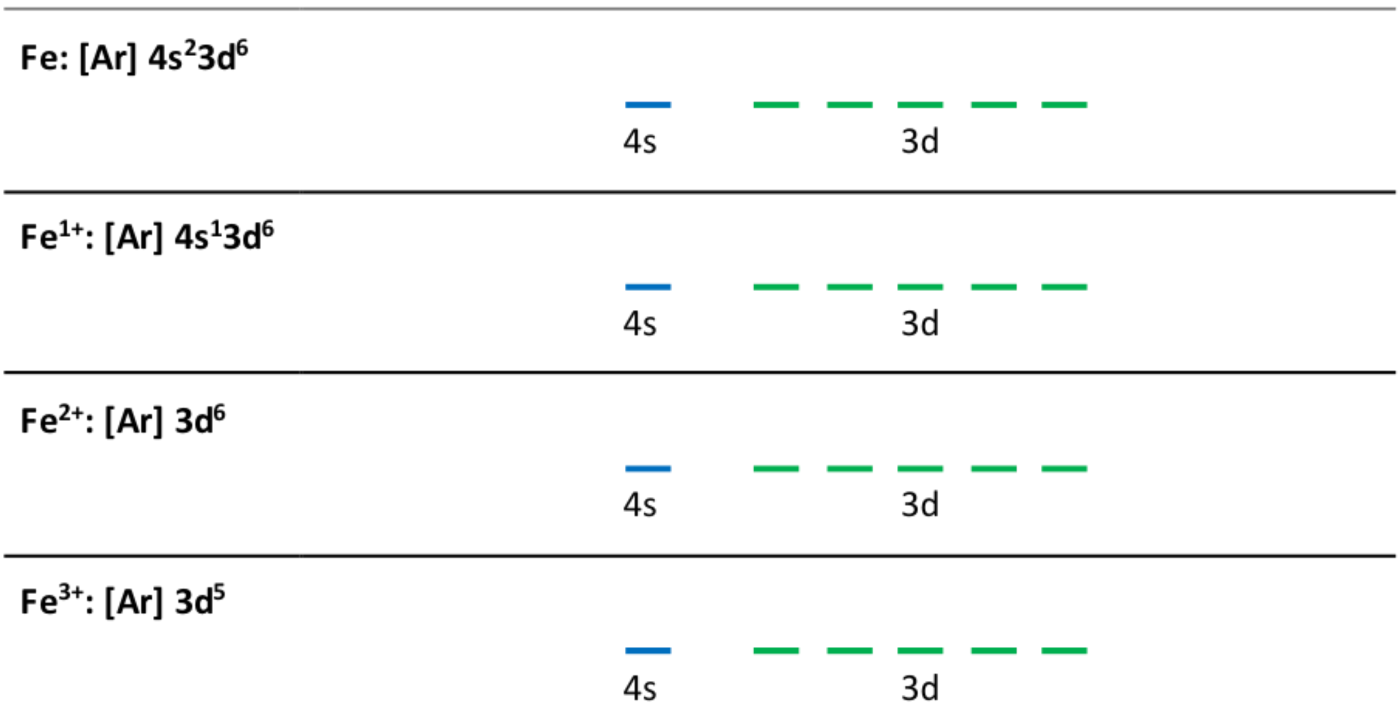
2.6.10

Electron Configuration of Ions of Transition Elements

The same rules that we just learned apply for transition elements as well. Note that valence electrons are only removed, never core electrons.

WATCH OUT!

Transition metals can lose both “n” and “n-1” valence electrons, but “n” electrons are always lost first! Let's take a look!



WIZE TIP

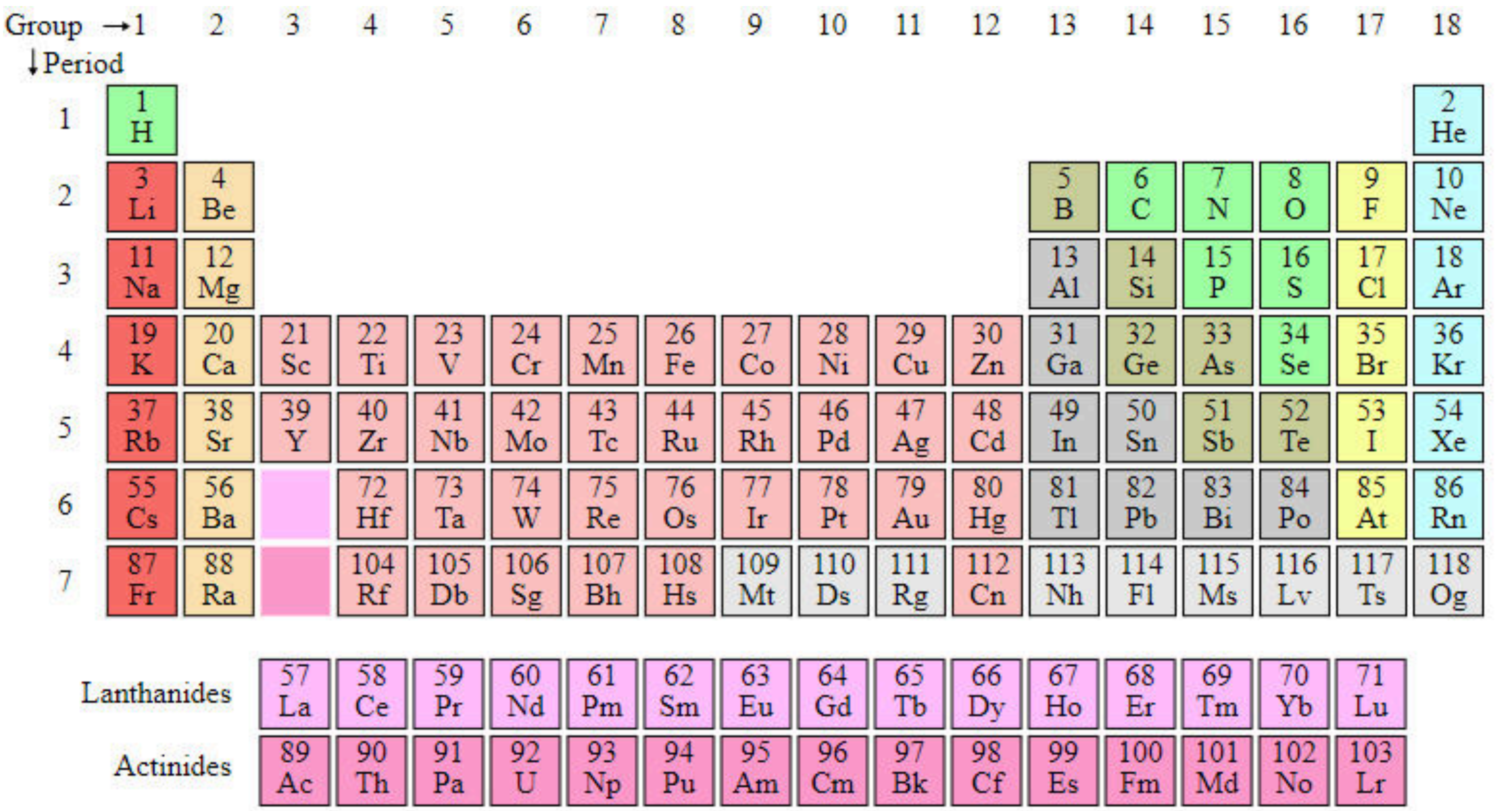
Always write out the **electron configuration for the neutral atom first.**
Then write out the **electron configuration for the charged element.**

This will help you avoid mistakes on the exam!

2.6.11

Example: Isoelectronic Species

Two atoms or ions are said to be isoelectronic if the electron configuration is the same for both species. Which of the following pairs are isoelectronic?



i) N^+ and C

ii) C and B^+

iii) Cl^- and Ar

2.6.12

Example: Writing Electron Configurations of Ions

Write the electronic configuration for the following species:

Group → / ↓Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba		72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra		104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Ms	116 Lv	117 Ts	118 Og
Lanthanides			57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu	
Actinides			89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr	

a) N^{+}

b) Ni^{2+}

2.6.13

Practice: Electron Configurations of Ions

For which of the atom/ions below, is the given electron configuration correct? (select all that apply)

	Species	Electron Configurations
i.	Cu	[Ar] $4s^2\ 3d^9$
ii.	Ti^{2+}	[Ar] $4s^2$
iii.	As^{5+}	[Ar] $4s^2\ 3d^8$

- i. ☐
- ii. ☐
- iii. ☐
- none of the above ☐

2.6.14

Diamagnetic vs Paramagnetic

You may see these terms come up when talking about electron configurations.

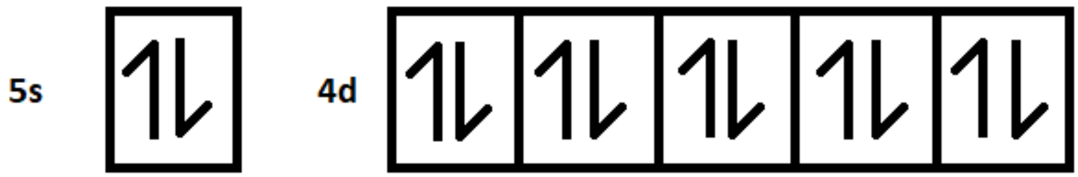
Diamagnetic - all electrons are **paired**

Paramagnetic - all electrons are **not paired**

Are the following electron configurations diamagnetic or paramagnetic?

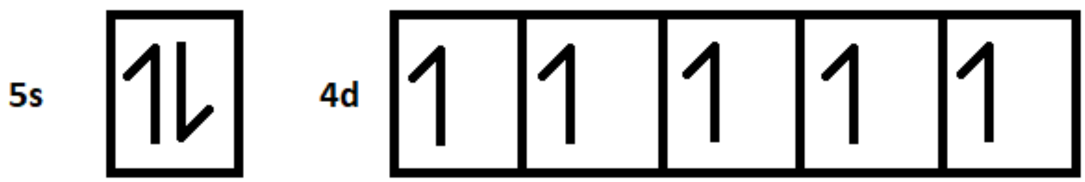
Cd

[Kr] $5s^2$ $4d^{10}$



Tc

[Kr] $5s^2$ $4d^5$



Group → ↓Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba		72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra		104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Ms	116 Lv	117 Ts	118 Og
Lanthanides			57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu	
Actinides			89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr	

Consider the electron configuration for Na

Is it diamagnetic or paramagnetic? ______________

Consider the electron configuration for Si

Is it diamagnetic or paramagnetic? ______________

Consider the electron configuration for Be

Is it diamagnetic or paramagnetic? ______________

WIZE CONCEPT

Diamagnetic-all electrons are **paired**. If something is diamagnetic it will also be **repelled** from externally produced magnetic fields.

Paramagnetic-One or more electrons are left **unpaired.** If something is paramagnetic it will be **attracted** to an externally produced magnetic field.

2.6.15

Example: Paramagnetic and Diamagnetic Species

Write the electronic configuration of the following species, and label them as **D** diamagnetic or **P** paramagnetic

Group → / ↓Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba		72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra		104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Ms	116 Lv	117 Ts	118 Og

Lanthanides	57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
Actinides	89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

a) Sc^{2+}

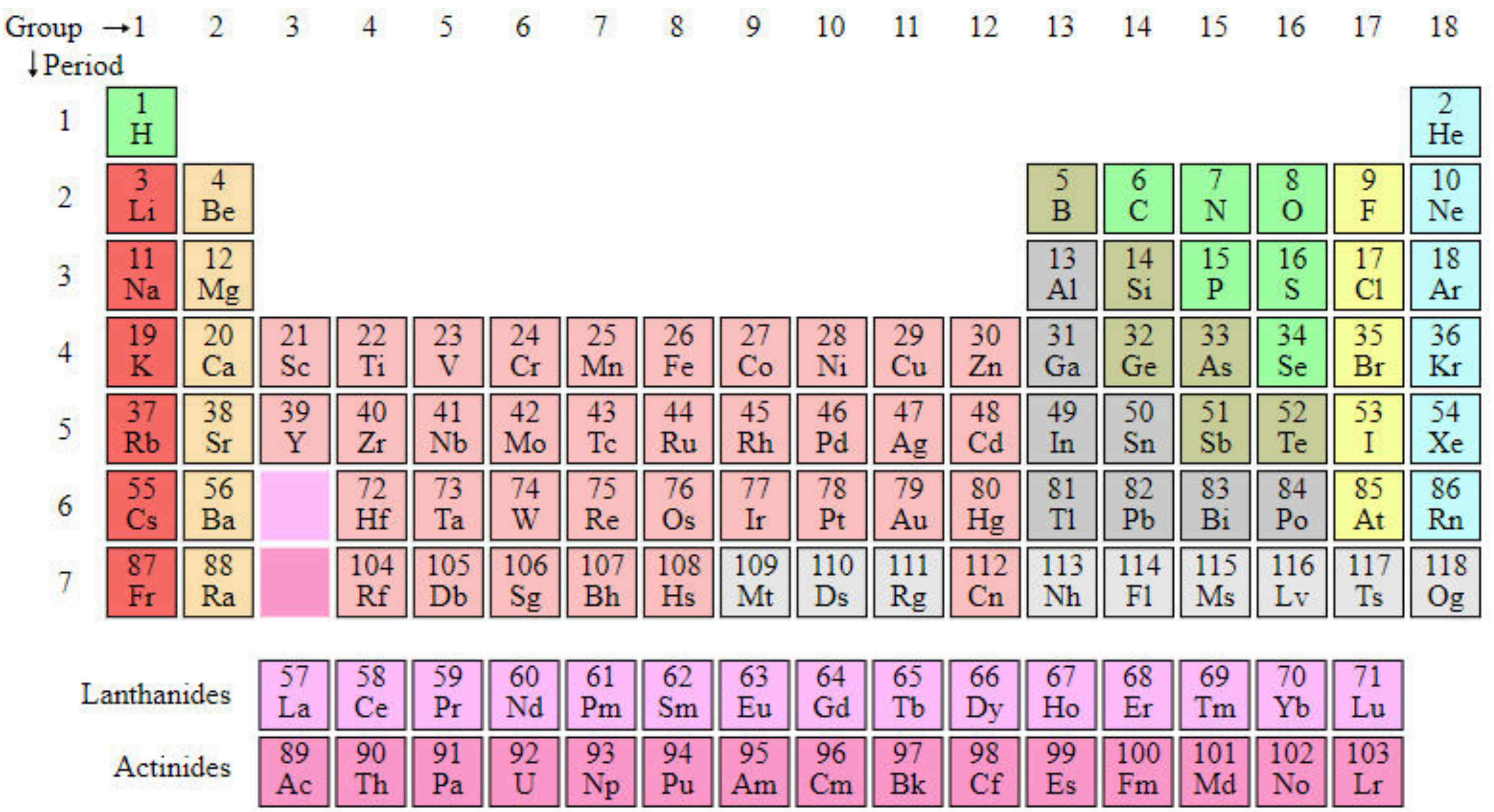
b) Cr^{2+}

c) Mn^{7+}

2.6.16

Practice: Diamagnetic

Which of the following are diamagnetic? (select all that apply in your answer)



- [ ] Cr
- [ ] Co
- [ ] Ru
- [ ] None of the above

View Solutions on Wizeprep.com

Solutions to these questions, as well as step-by-step breakdowns of the answers at:

2.6.17

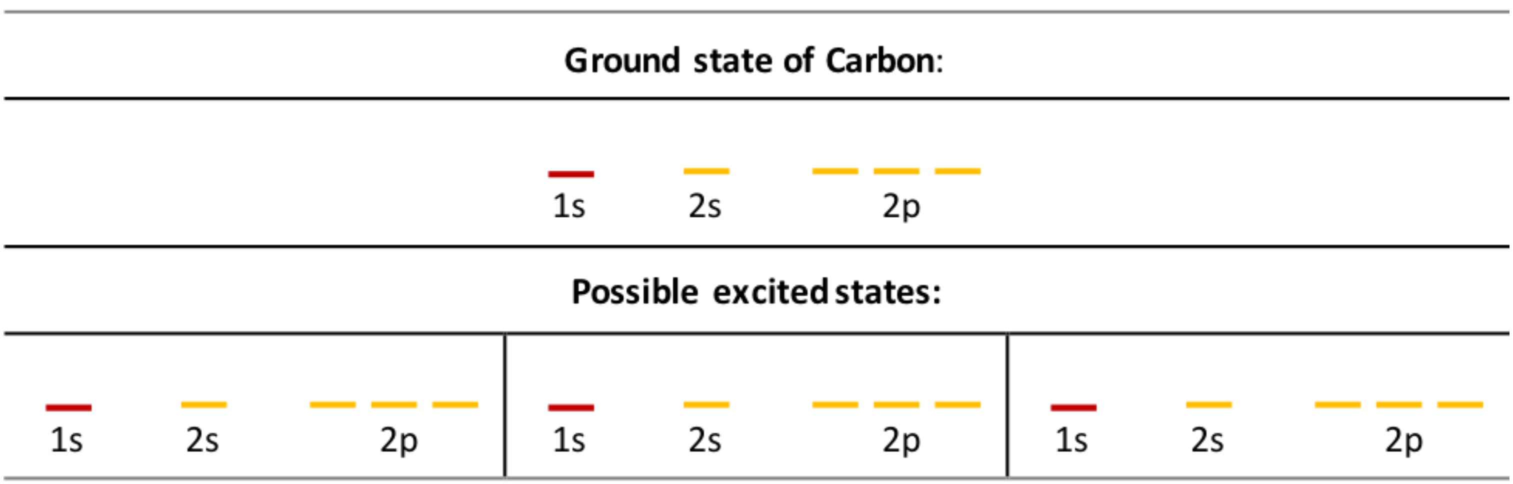
Ground State Vs Excited State Electron Configurations

Generally, when we find electron(s) in a higher energy level than expected for ground state, the atom is in an excited state.

WATCH OUT!

Do not confuse excited state configurations with those of ions.

Excited states have the **same number of electrons as the ground state atom**, whereas an ionic configuration will have a different a number of electrons than the atom it came from.



WIZE TIP

For excited states, the Aufbau Principle and Hund's Rule can be disobeyed, but the Pauli Exclusion Principle must ALWAYS be followed.

2.6.18

Example: Excited State Electronic Configurations

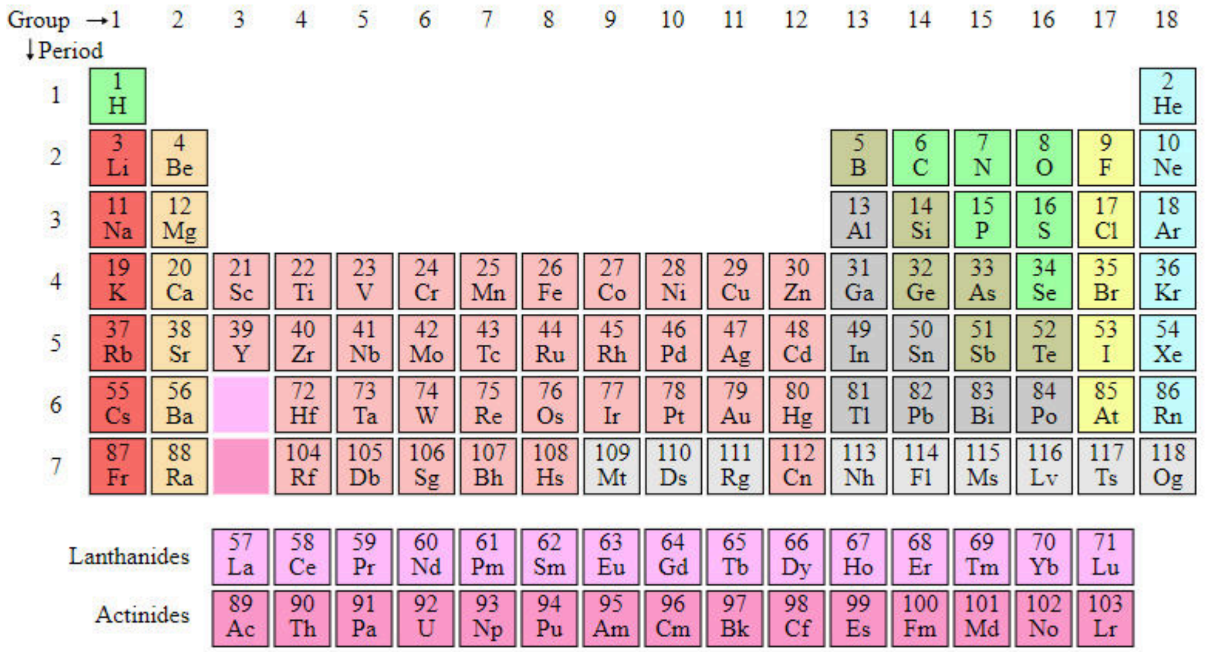
Determine the electronic configuration for a chlorine atom that has one electron excited from the 3p orbital to the 4s orbital.

Group → / ↓Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba		72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra		104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Ms	116 Lv	117 Ts	118 Og

Lanthanides	57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
Actinides	89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

Example: Excited States

How many electrons would an excited atom of chlorine have?



A. 6
B. 7
C. 16
D. 18
E. None of the above

2.6.20

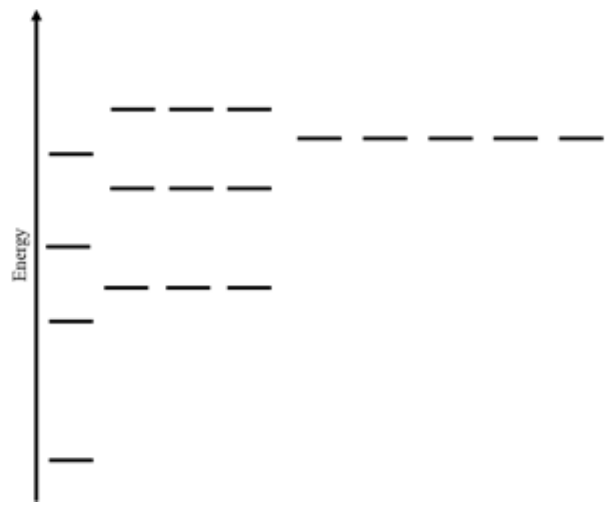
MARK YOURSELF QUESTION

1. Grab a piece of paper and try this problem yourself.
2. When you're done, check the "I have answered this question" box below.
3. View the solution and report whether you got it right or wrong.

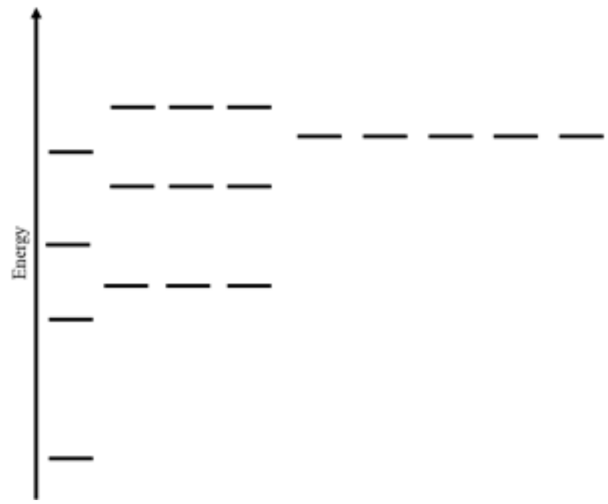
Practice: Excited State in an Energy Level Diagram

Consider an atom of the alkali metal Potassium.

a) In the energy diagram given, fill in the required number of electrons for a potassium atom and name the occupied orbitals.



b) The valence electron of potassium can be excited to the next highest energy state in the same valence shell. Below, show how the energy level diagram for the excited atom of potassium has changed. Make sure to name all occupied orbitals.


Energy

Practice: Ground or Excited State

Identify the following atom or ion and decide whether it is in a ground or excited state

Z = 26 electronic configuration: $[Ar]4s^2 3d^3$

- Fe^{3+} Excited state
- Fe^{3+} Ground state
- Fe^{2+} Excited state
- Fe^{2+} Ground state