

Queen's **CHEM 112A**

Fall 2025, Chapter 3 Notes



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3. Atoms

3.1 Electromagnetic Radiation

3.1.1

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

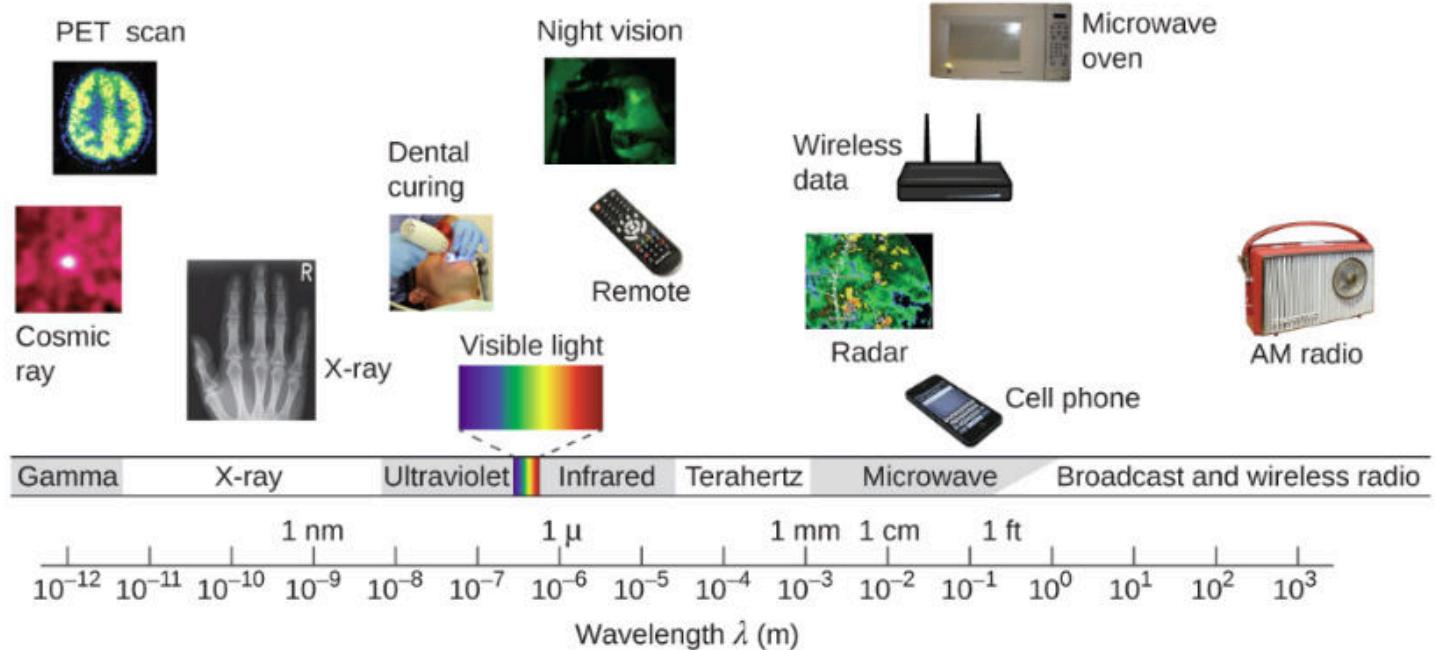


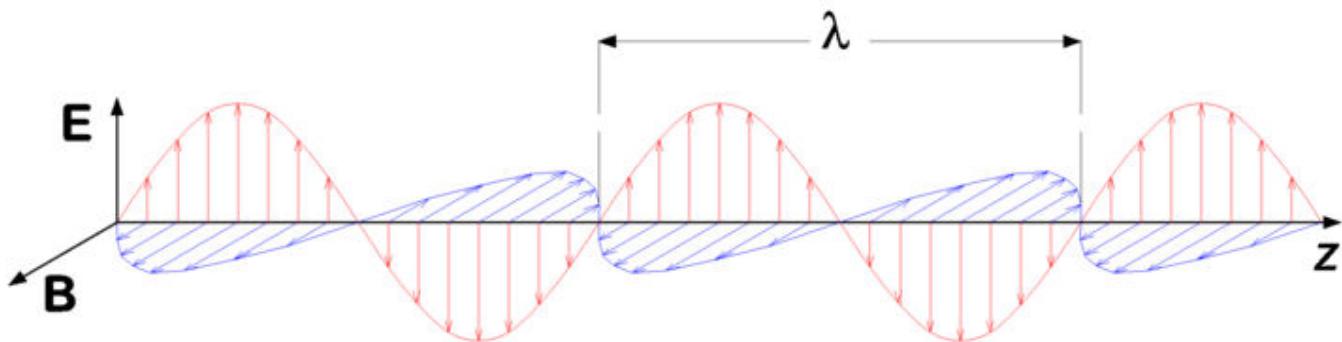
Photo by Rice University/ CC BY

- 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



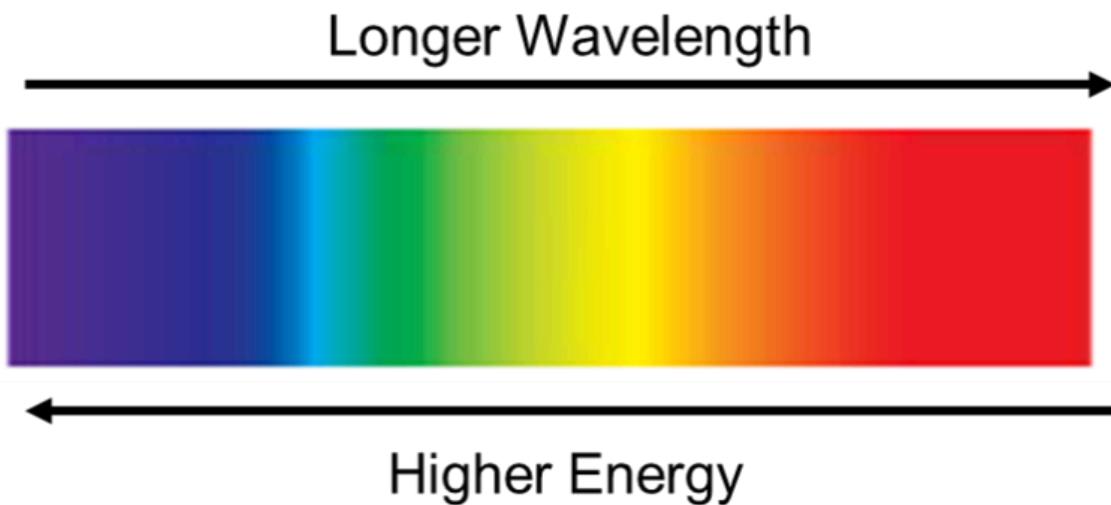
where E is the electric field and B is the magnetic field

- 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

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

where ***h*** is Planck's constant: 6.626×10^{-34} Js
and ***c*** is the speed of light: 3.0×10^8 m/s



i WIZE TIP

You should memorize the value for the speed of light (*c*) = 3.0×10^8 m/s as that value is not always provided on exams.

💡 WIZE CONCEPT

Where will I see this on an exam?

Questions on this topic tend to arise in the form of a calculation. You will be asked to determine the energy or wavelength of a given ray of light.

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

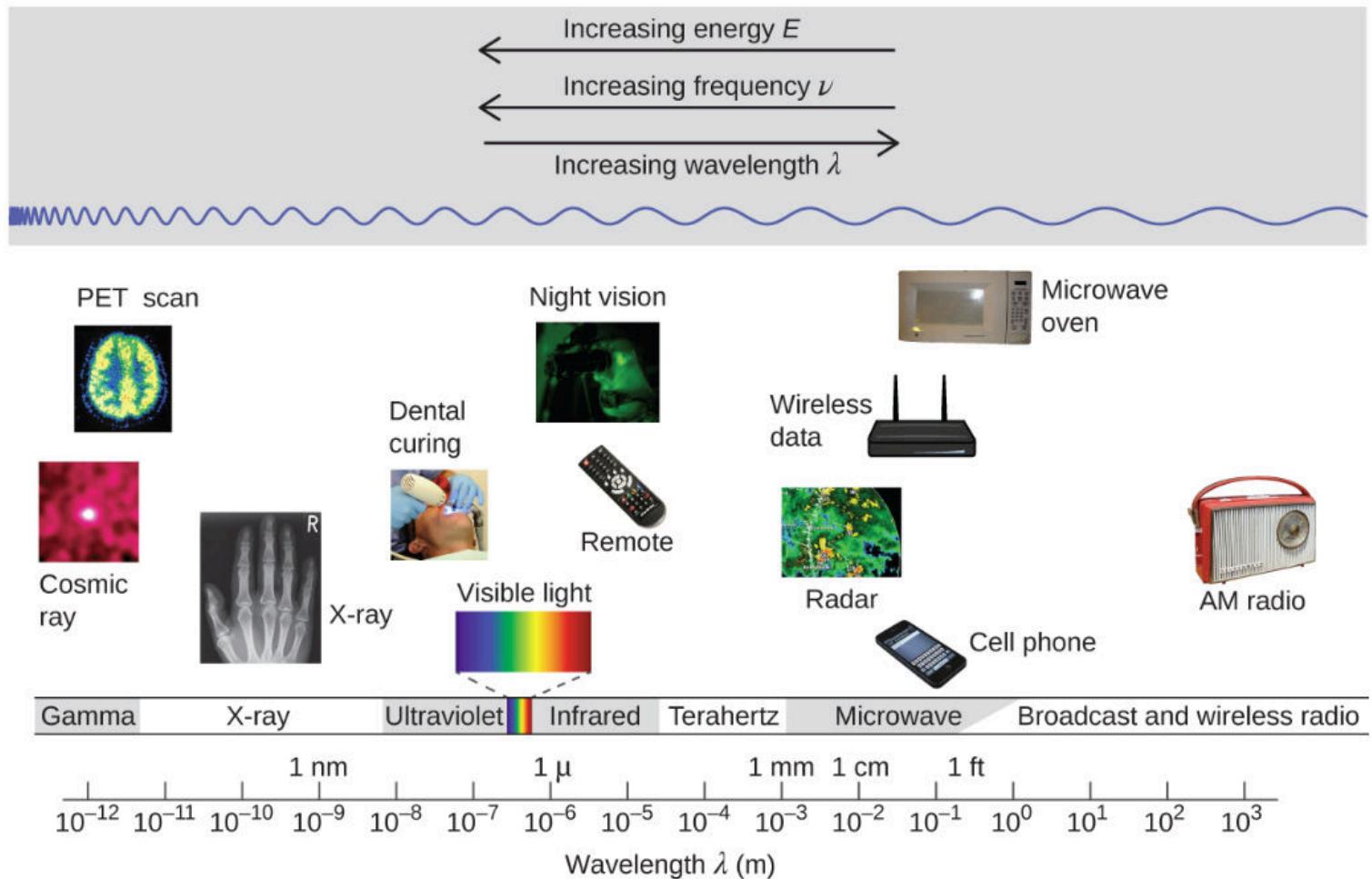


Photo by Rice University/ CC BY

3.1.4 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?
 $h=6.626 \times 10^{-34} \text{ Js}$

! WATCH OUT!

Don't forget that for the photon equations **E must be in joules (J) and λ (wavelength) must be in m (meters) not nm (nanometers)!**

- If you are given a wavelength value in nm, you would have to convert it to m before plugging it into an equation.

Remember: $1\text{nm}=1\times10^{-9}\text{m}$

Example: Calculating Frequency of a Photon Given Energy

Calculate the frequency of a photon with an energy of 236 kJ.

$$h=6.626 \times 10^{-34} \text{ Js}$$

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

3.1.7

Practice: Calculating Frequency of a Photon Given Wavelength

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

$4.29 \times 10^{-11} \text{ s}^{-1}$



$6.72 \times 10^{12} \text{ s}^{-1}$



$4.29 \times 10^{14} \text{ s}^{-1}$



$7.23 \times 10^{16} \text{ s}^{-1}$



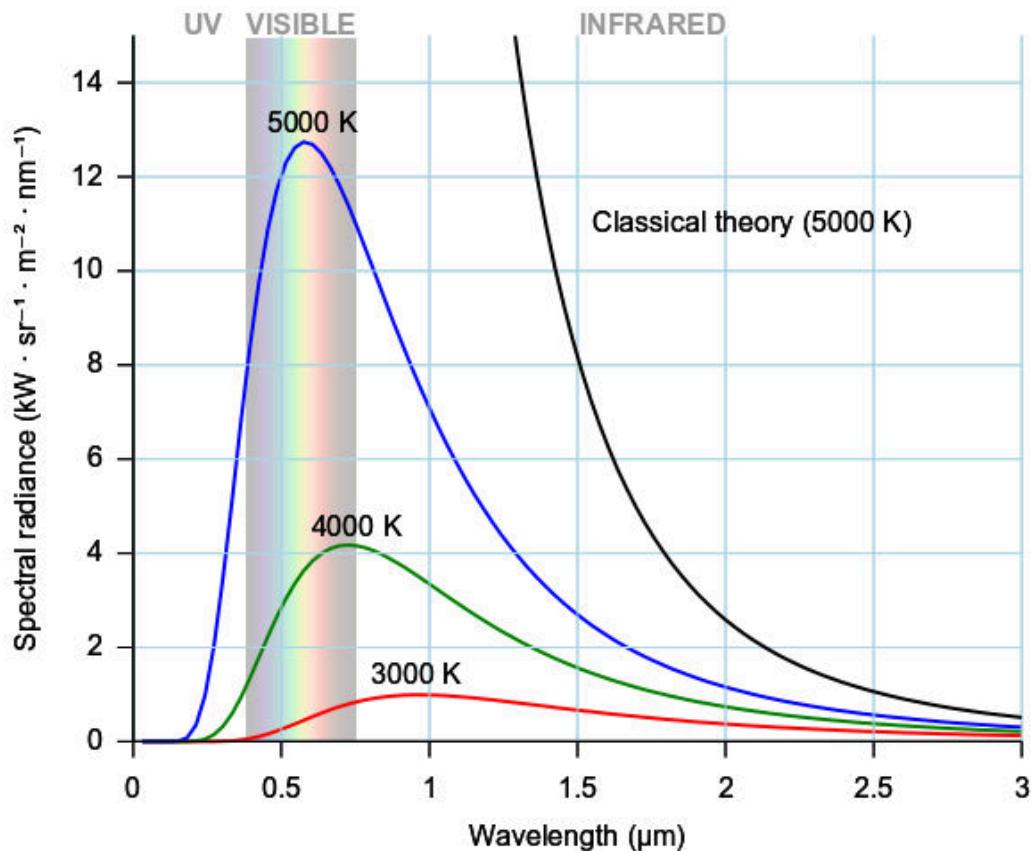
3.2

Blackbody Radiation

3.2.1

Blackbody radiation

Classical Problem:



Issues:

1. Classical Theory does not do a good job predict the emission of bodies in the infrared region, especially for cooler bodies.
2. Near the UV-range classical theory predicts an infinite increase in radiance which means an infinite amount of light being emitted. This is a clear contradiction of energy conservation, a BIG ISSUE! Known as the UV-Catastrophe

Solution:

- Max Plank discovered that if he restricted energy that can be released or absorbed by atoms into discrete energy packets he could successfully predict the experimental observations.

Results:

- The smallest quantity of energy that can be emitted or absorbed as electromagnetic radiation, is called a quanta

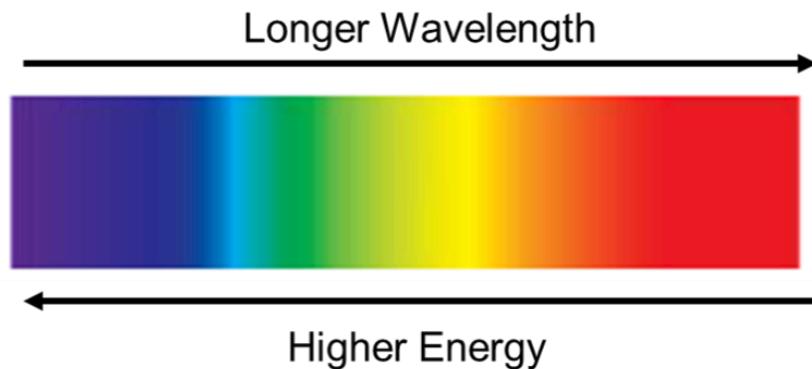
The energy of a single quantum is:

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

E = energy of the photon in J

$h = 6.626 \times 10^{-34} \text{ Js}$ (Planck's Constant)

ν = frequency in s^{-1}



i WIZE TIP

You should memorize the value for the speed of light (c) = $3.0 \times 10^8 \text{ m/s}$ as that value is not always provided on exams.

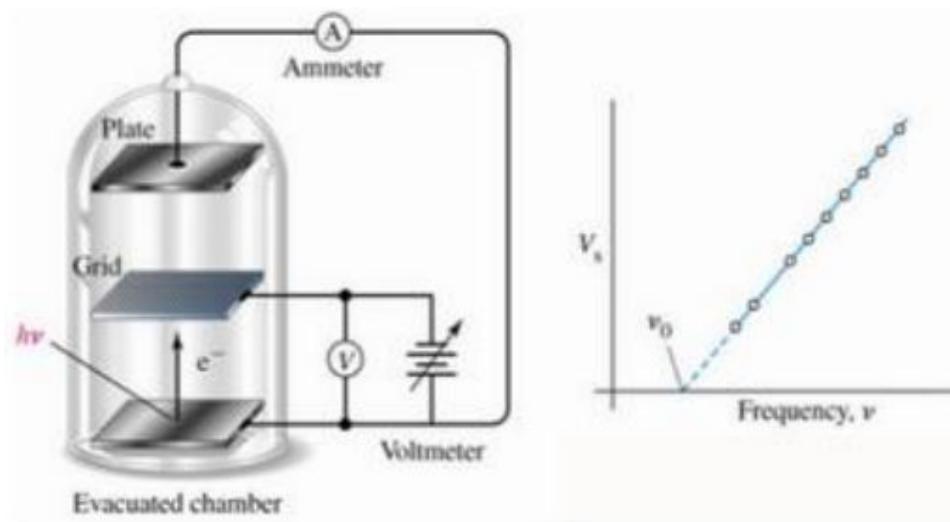
3.3

The Photoelectric Effect

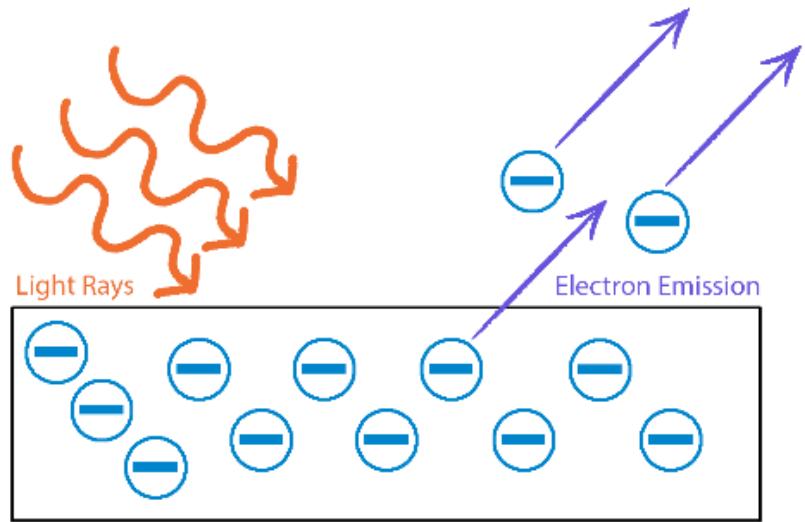
3.3.1

The Photoelectric Effect

In 1905, Einstein conducted his **photoelectric effect experiment**.

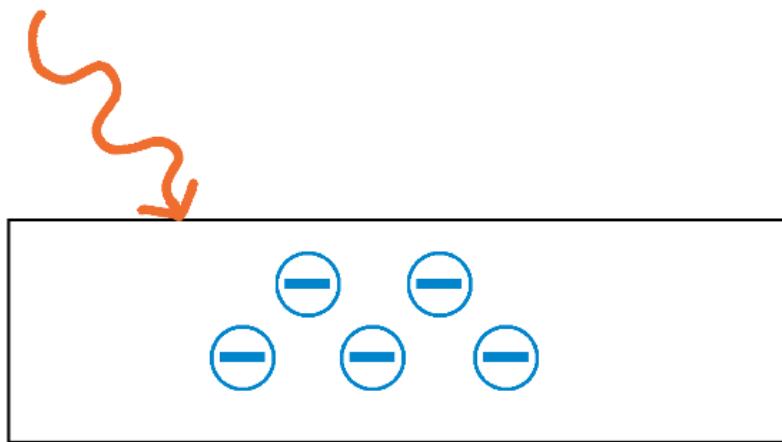


- The apparatus is made of a metal grid in which light of different intensities and frequencies was shown onto it.
- The ammeter (A) measures current (the **number of electrons emitted**).
- The voltmeter (V) measures the stopping potential (the **energy of the electrons emitted**).

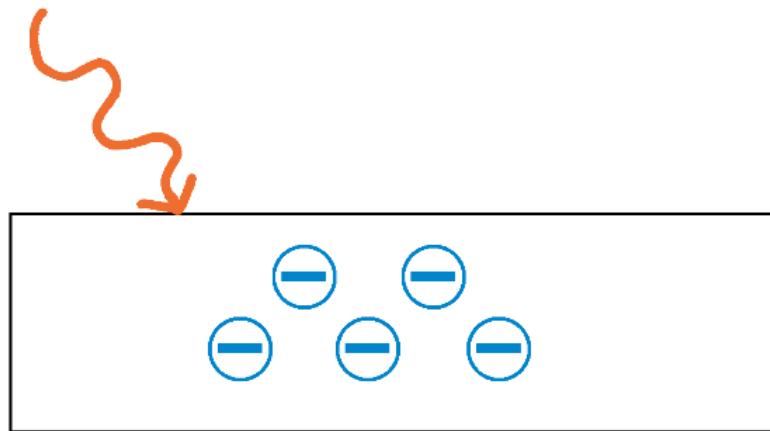


Experimental Findings

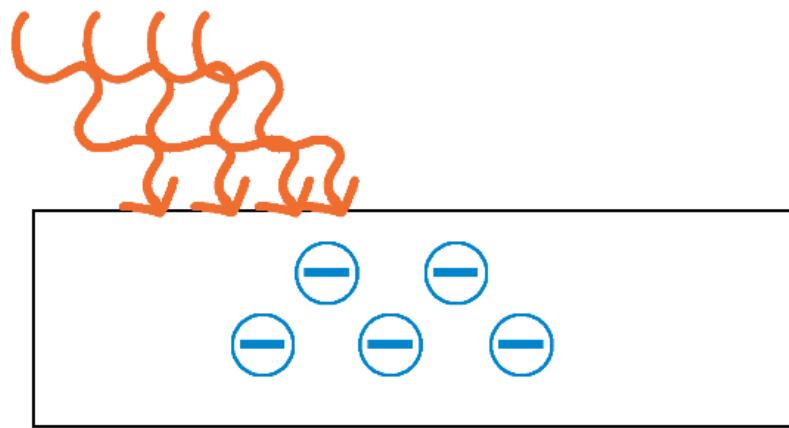
1) 1 photon with **sufficient energy** to eject an electron:



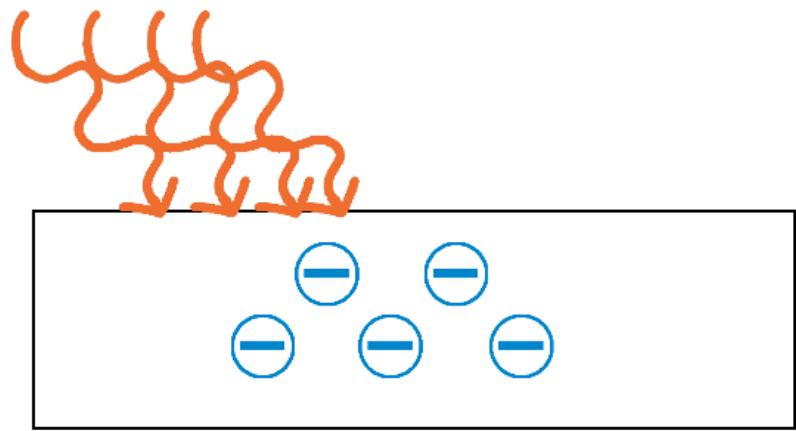
2) 1 photon with **insufficient energy** to eject an electron



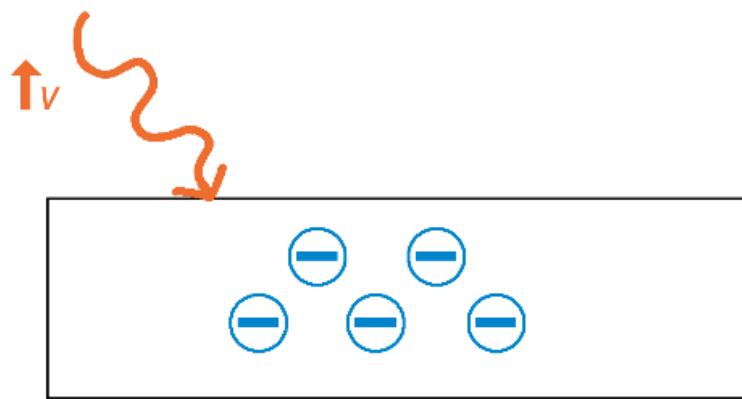
3) Multiple photons hitting the metal, **individually they have insufficient energy** to eject an electron but if you add up the their energy they have more energy than required to eject an electron



4) Multiple photons, each with sufficient energy needed to eject an electron hitting the metal



5) A photon with an **increased frequency** than the photons with minimum sufficient energy to eject an electron



Summary of the Findings:

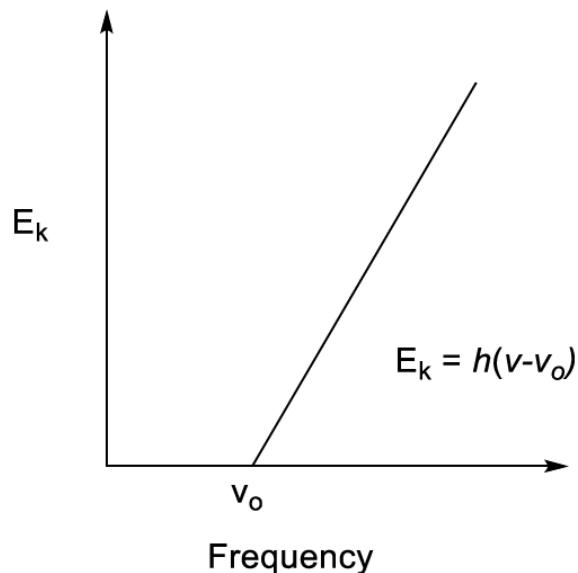
- Electrons must be ejected by 1:1, photon : electron collisions
- Increased intensity of the light (more photons for a given part of the light beam), increased the number of electrons emitted (if each of the photons had sufficient energy to eject electrons)
- Increased frequency of the light, increased the energy of the electrons emitted

Before this experiment, people thought of light only as WAVES, but Einstein showed that light could also be thought of as particles (photons).

- When particle A hits particle B, particle A transfers its momentum to particle B. This is not the case for waves!
- The photoelectric effect showed that light can behave like particles since photons of light can hit a solid substance and cause electrons to be emitted

Photoelectric Effect Equation

This experiment shows that light can behave like a particle “knocking” electrons out of a material.



- The **y-axis** shows the **kinetic energy (E_k) of the emitted electron**
- The **x-axis** shows the **frequency of the photon** that is hitting the metal's surface
 - When the **frequency is too low**, the energy of the photon is too low and an **electron is not emitted**
 - When the **frequency reaches the threshold frequency**, there is **enough energy to emit an electron**
 - At even **higher frequencies**, the **kinetic energy (E_k) of the electron is increased further!**

The energy corresponding to v_0 is called the '**binding energy**' and represents the amount of **energy required to remove the electron, E_{Binding}**

$$E_k \text{ electron being ejected} = E_{\text{photon}} - E_{\text{binding}}$$

$$\text{where } E_{\text{photon}} = h\nu = \frac{hc}{\lambda}$$

$$\text{and } E_{k_{\text{electron}}} = \frac{1}{2}mv^2$$

and E_{binding} =minimum amount of energy needed to eject an electron

 WIZE TIP

The **binding energy for each metal is unique**.

Example: The binding energy of iron and gold are different.

The **binding energy for a specific metal is constant** so if you are given the binding energy for iron in a question, you can use that for part a) b) c) for example.

Does an electron get displaced when $E_{\text{photon}} >$ or $<$ than E_{binding} ?

Example: Ejecting an Electron

When lithium metal is irradiated by a laser with a wavelength of 400 nm no electrons are ejected from the surface. If you want to ionize lithium which of the following types of radiation should you use?

- a) UV-Laser
- b) IR-Radiation
- c) Red Light Laser
- d) Green Light Laser
- e) Radio Wave

 **WIZE TIP**

When you see a question that mentions **electrons being ejected** from the surface of a metal, you know the question is going to have something to do with the **photoelectric effect!**

Practice: No Electrons are Ejected

A metal surface is irradiated by light from a laser with a wavelength of 492 nm. No electrons are measured leaving the surface. What is the most likely explanation?

All of the electrons which are easy to remove have already been removed

The frequency of the photons is less than the threshold frequency of the metal

The energy of the photons is greater than the threshold value

The light is not irradiating the surface at the correct angle

None of the above

3.3.5

Practice: Calculate the Binding Energy

An element is bombarded with a beam of X-rays of wavelength 0.75664nm. Upon irradiation, electrons are ejected with a velocity of 2.35×10^7 m/s. The mass of an electron is 9.11×10^{-31} kg.

What is the binding energy for the element?

$$h=6.626 \times 10^{-34} \text{ Js}$$

$$1.12 \times 10^{-14} J$$

$$2.46 \times 10^{-17} J$$

$$1.12 \times 10^{-17} J$$

$$0 J$$

3.4

Understanding the Bohr Model

3.4.1

The Basics of Absorption and Emission Spectra

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

Continuous Spectrum



Emission Lines



Absorption Lines



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

Absorption Lines



Emission Spectra

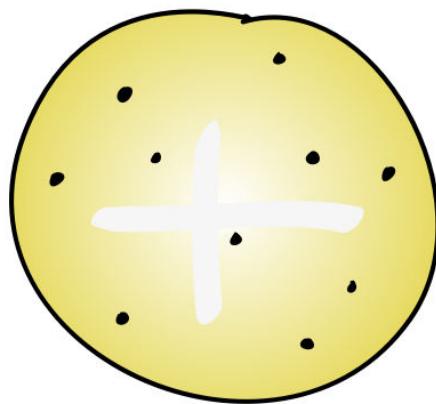
- We get an emission spectrum by measuring the electromagnetic radiation (light) that is emitted from a substance
- The background is black and the **coloured bands** represent the **wavelengths of photons that were emitted** from this atom's electrons

Emission Lines



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

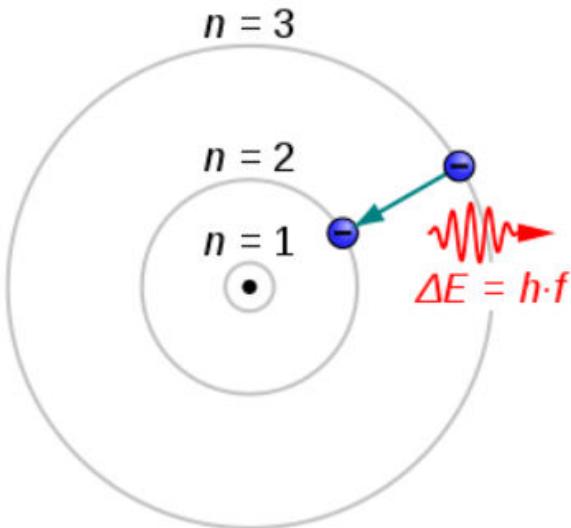


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- Electrons orbit nucleus at **fixed distances** and in **circular paths** (like how planets orbit the sun!)
- Energy electrons have are **quantized** (only specific energies are allowed)
- Shell's distance from nucleus determines energy ($n=1, 2, 3$, 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!

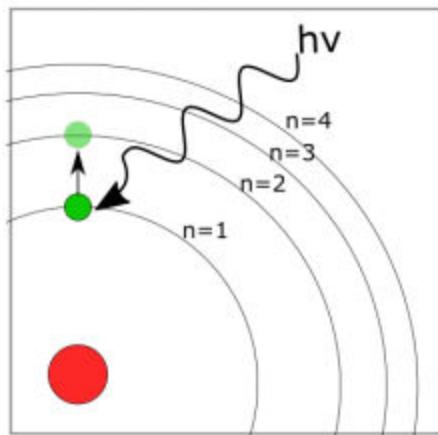
Analogy: Staircase vs Ramp



- **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)

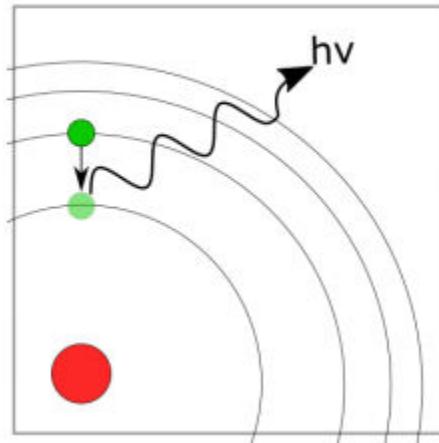
Absorption vs Emission of a Photon

Absorption



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

Emission



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

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

Example: Electron Transitions

An electron in the $n = 4$ shell of an excited hydrogen atom relaxes to its 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 $n = 1$

3.5 Bohr Model Equations

3.5.1 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

$$E = \frac{(-2.178 \times 10^{-18} \text{ J}) (Z^2)}{n^2}$$

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

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

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= $1.097 \times 10^7 m^{-1}$

n_i =initial energy level, **n_f** =final energy level

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 $n = 3$?

3.5.4

Practice: Using Bohr Model Equations to Determine an Energy Level

A photon of light was released by a hydrogen atom with an energy of 1.94×10^{-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

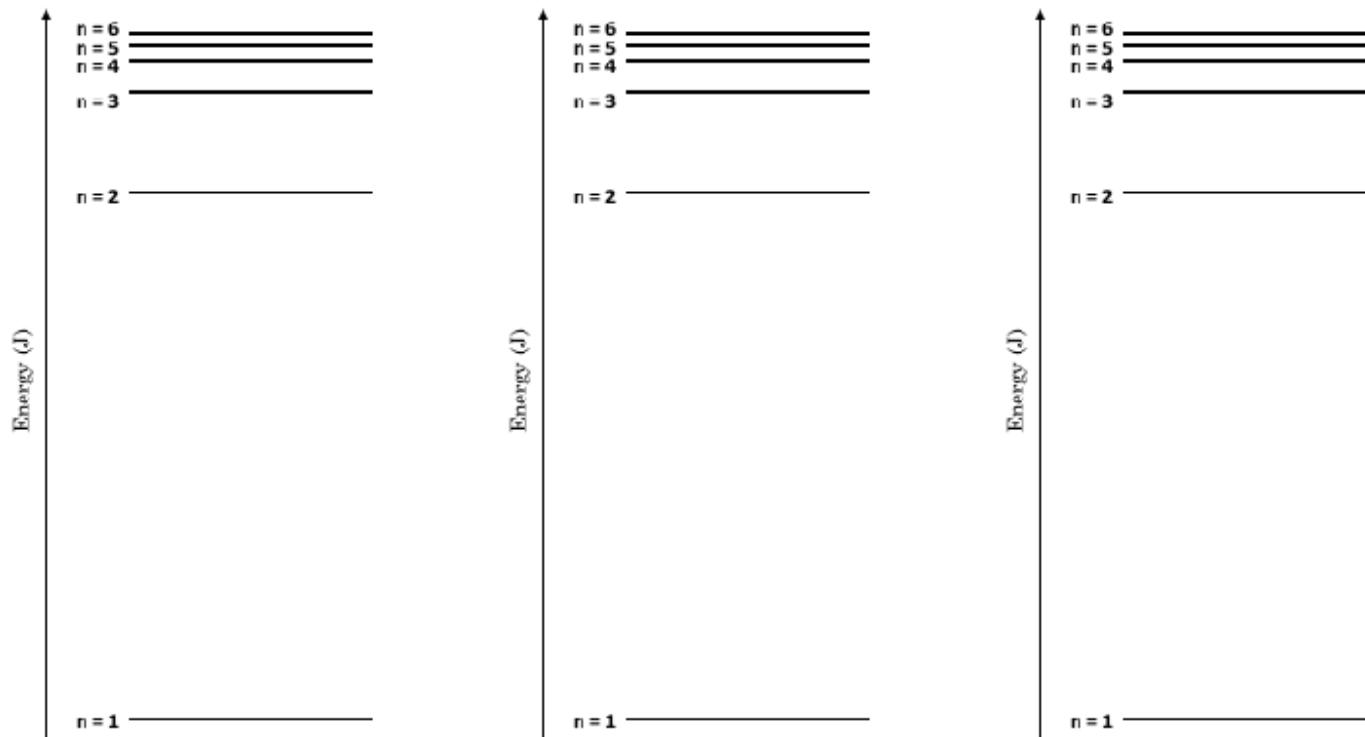
3.6

How Atomic Spectra Relates to Electronic Transitions + Other Applications of the Bohr Model

3.6.1

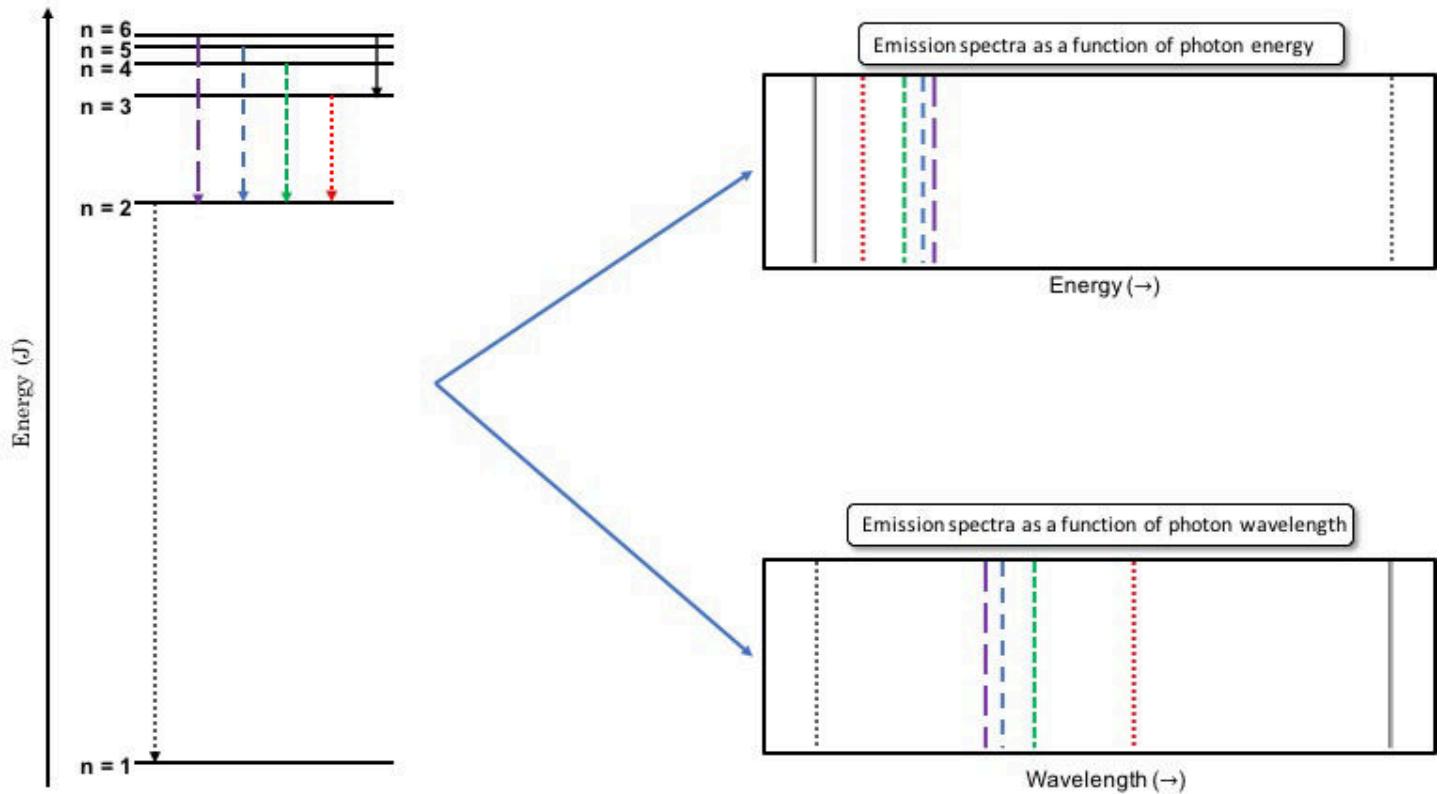
Atomic Line Spectra

Emission and Absorption



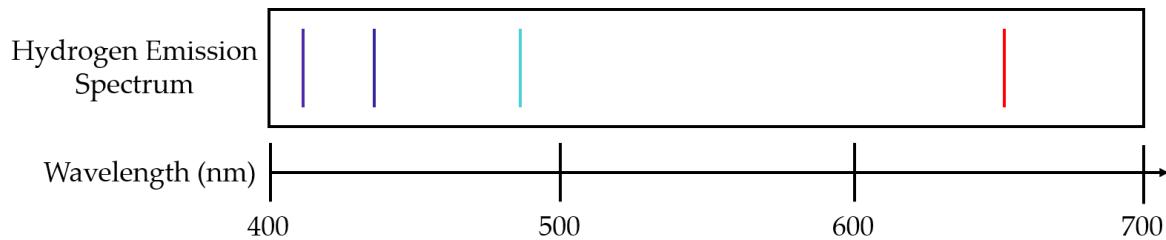
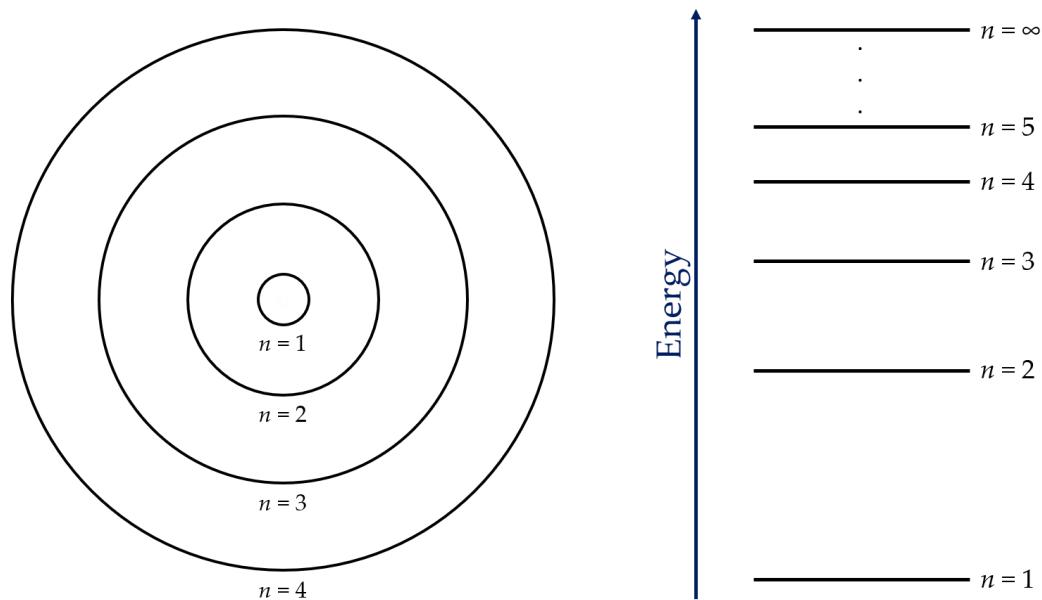
Note: **n=1 is ground state**, all other n levels are excited states

Emission Spectra



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 4th 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:

3.6.4

Practice: Understanding Emissions

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

n=2 to n=1

n=1 to n=6

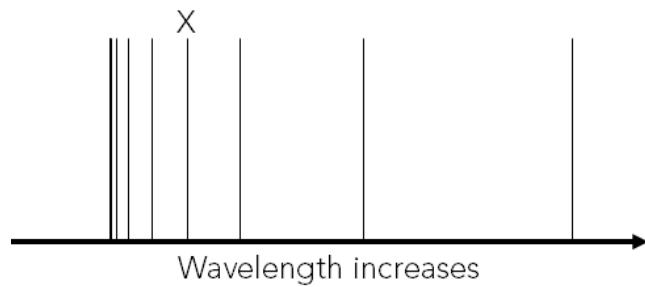
n=5 to n=1

n=7 to n=2

n=1 to n=4

3.6.5

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 $n = 1$

3.7 Standing Waves

3.7.1

Standing Waves & Particle in a 1-D Box

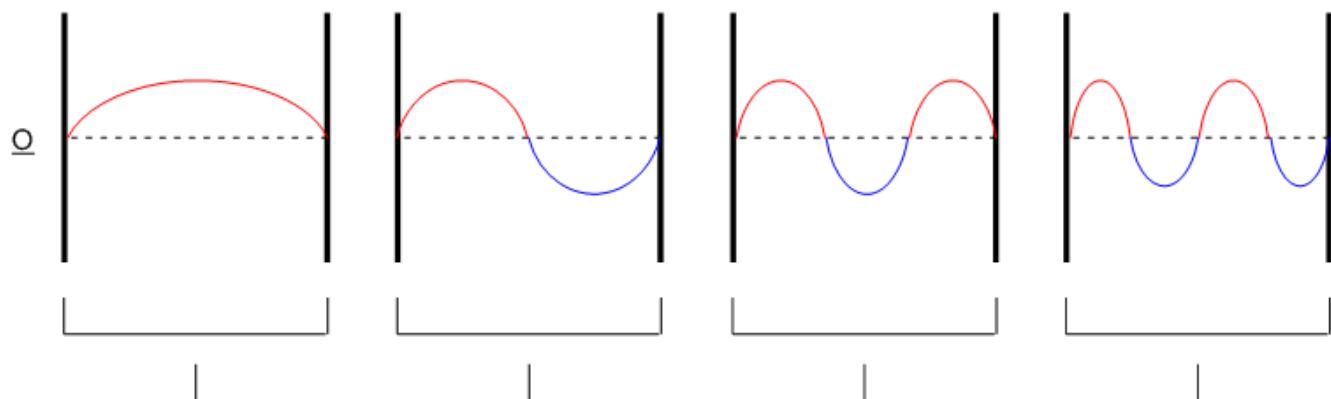
Standing Waves

Definition: A wave which is confined between two locations in space.

Node: The (non-end) points where a wave undergoes no displacement.

- Because electrons display wave-like characteristics, they can be approximated as standing waves.
- Examining the outcomes from a 1-D particle in a box will serve as an illustration for the 3-D “boxes” that electrons occupy, called orbitals.
- The confinement, as shown in figure 3.4, necessitates that only certain wavelengths are permitted.

$$\lambda = (2L) \frac{1}{n}$$



\underline{L}

$\psi_1(x)$
 $n = 1$
0 nodes
 $\lambda = 2L$

\underline{L}

$\psi_2(x)$
 $n = 2$
1 node
 $\lambda = L$

\underline{L}

$\psi_3(x)$
 $n = 3$
2 nodes
 $\lambda = 2L/3$

\underline{L}

$\psi_4(x)$
 $n = 4$
3 nodes
 $\lambda = L/2$

Wavefunctions

- The standing waves shown above can be described by mathematical functions, or, wavefunctions

$$\psi_n(x) = \sqrt{\frac{2}{L}} \sin\left(\frac{n\pi x}{L}\right)$$

- The sin wave provides the oscillation between positive and negative values
 - L is the length of the box
 - The $\sqrt{\frac{2}{L}}$ term is a normalization term and you are NOT required to understand what it does (for the purpose of this course).
-
- The probability of finding a particle at a given position is found by squaring the wavefunction and is denoted ψ^2

Kinetic Energy and Quantization

- When we think about kinetic energy in the macroscopic world, it seems fairly intuitive that kinetic energy could be any value and varies linearly with velocity
- In the world of subatomic particles, the quantization of wavelength results in a quantization of other properties like kinetic energy
- Using the classical equation for kinetic energy and momentum, we can add in the equation of de Broglie as well as the wavelength of a particle in one-dimensional box. This allows us to express kinetic energy in terms of n and L as shown below:

$$E_k = \frac{n^2 h^2}{8mL^2}$$

- n represents which wavefunction we are looking at ($n=1,2,3,4\dots$)
- m is the mass of the particle, typically an electron.
- L is the length of the box
- Note that this is the equation for kinetic energy of a particle in a 1-D box

3.7.2

What is the wavelength of a $n = 4$ standing wave confined to a 210nm length?

3.7.3

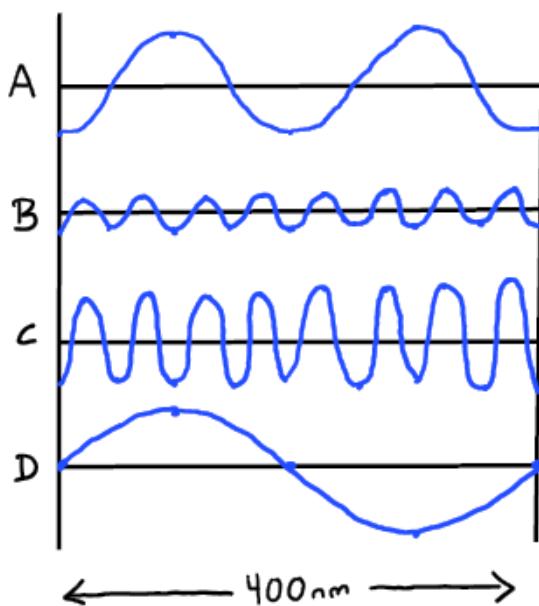
Complete the following sentence with the choice that best describes a standing wave

A standing wave is a wave where,

- a) The nodes do not move along the length of the wave
- b) The amplitude of the wave is the same at each maxima
- c) The end-points are stationary
- d) An electron is in the wave
- e) The wave travels above the speed of light

3.7.4

Use the graph below to answer the following questions:



1. Which wave has the lowest energy? _____

Solution:

2. What is the wavelength of wave A? _____

Solution:

3. If wave B represents a photon from the UV region, what region of the electromagnetic spectrum would you find wave C? _____

Solution:

4. Which wave has the lowest amplitude? _____

Solution:

3.8 Wave-Particle Duality & the de Broglie Equation

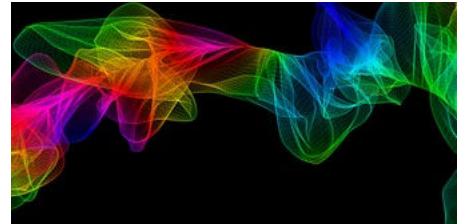
3.8.1

Wave-Particle Duality

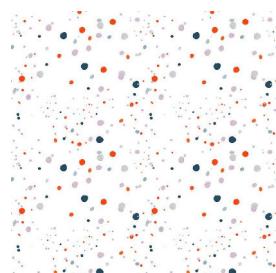
Light was originally thought of as only having wave-like properties, but was later shown to have particle-like properties as well.

Light as Waves

- When light was shone 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:

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

De Broglie Equation

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

h is **Planck's constant**= 6.626×10^{-34} Js (or 6.626×10^{-34} kg m²/s)

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

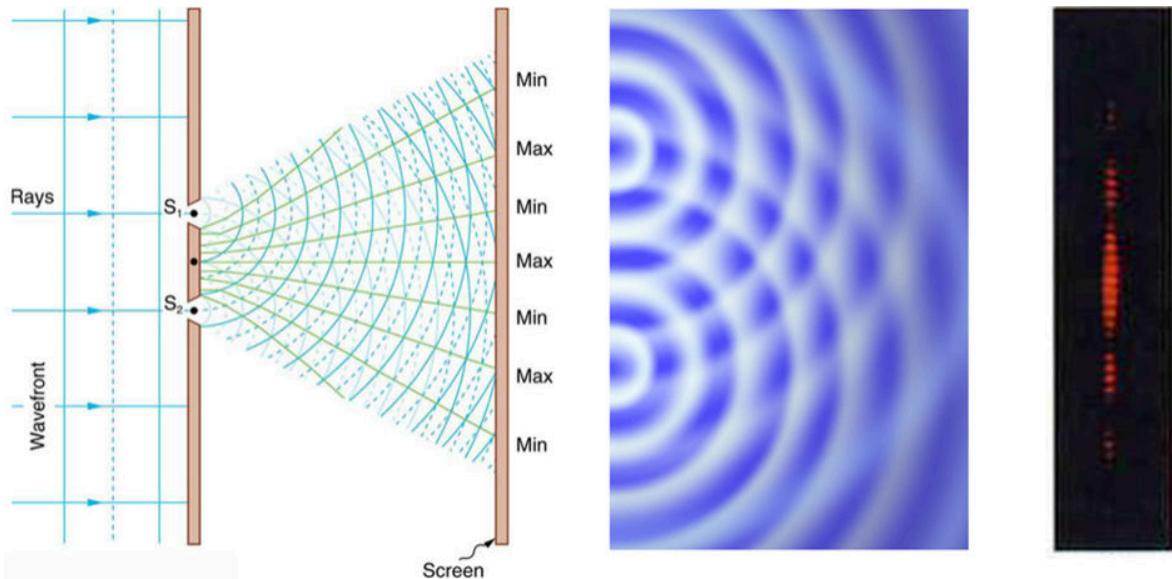
v is the velocity of the particle (**m/s**)

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

Photo by Rice University / CC BY

Could electrons also act as waves?

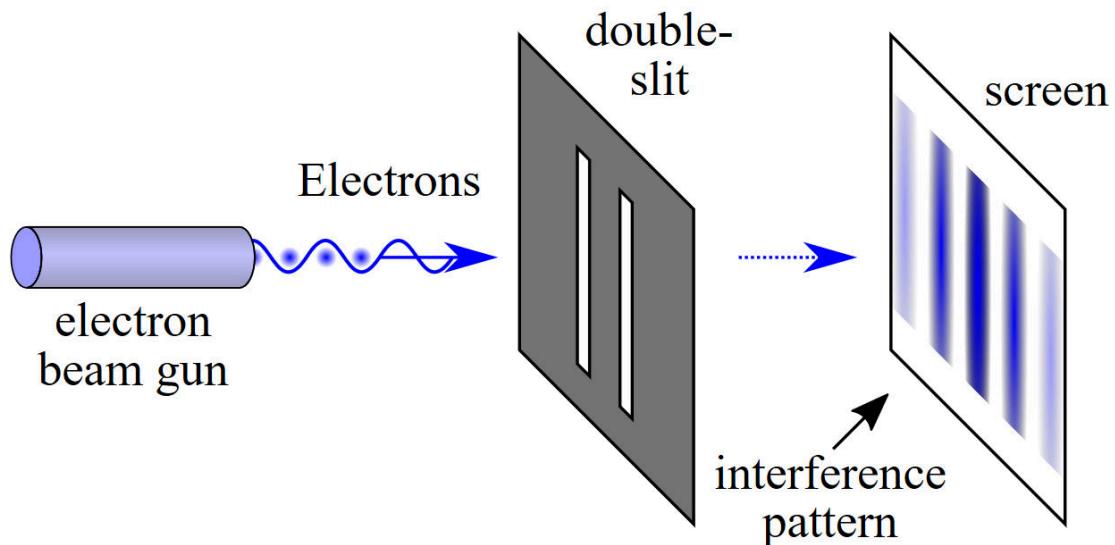


Photo by Johannes Kalliauer / CC BY

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 2.4×10^8 m/s in a particle accelerator. The mass of the neutron is 1.675×10^{-27} kg, what is its wavelength? (h is 6.626×10^{-34} kg m²/s)

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

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?



3.9

Introduction to the Quantum Model

3.9.1

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

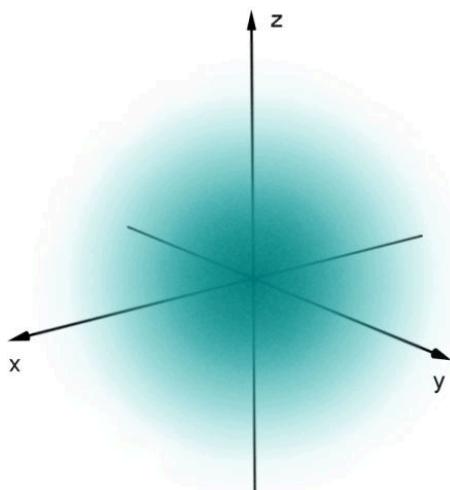
Couldn't explain how atoms were bound together

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.



Orbital s ($\ell = 0, m_\ell = 0$)

Electron Density Diagram

Photo by RJHall / CC BY

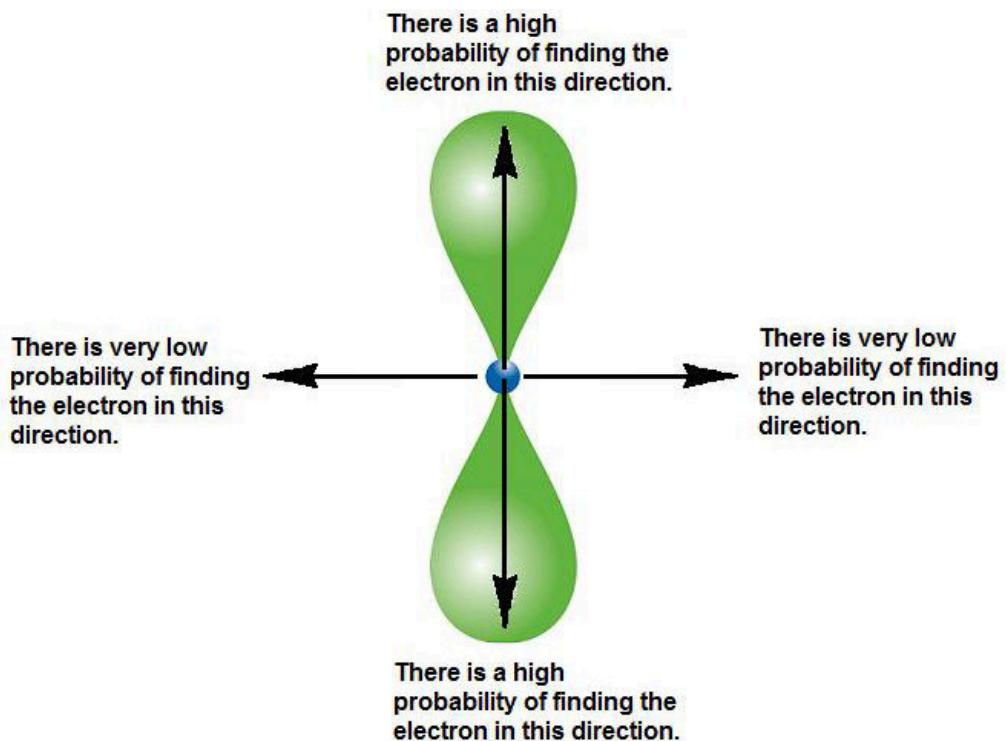


Photo by CK-12 Foundation / CC BY

3.10 Heisenberg Uncertainty Principal

3.10.1

The Heisenberg Uncertainty Principle

As electrons were discovered to have both wave and particle properties, Heisenberg found that we **cannot determine the exact position and momentum (direction and speed) simultaneously**. This is known as the Heisenberg uncertainty principle.

$$\Delta x \Delta p \geq \frac{h}{4\pi}$$

Δx is the **uncertainty in position**

Δp is the **uncertainty in momentum**

$\Delta p = m \Delta v$ where **m** is the mass (**kg**) and **Δv** is the uncertainty in velocity (**m/s**)

h is Planck's constant ($6.626 \times 10^{-34} \text{ Js}$ or $6.626 \times 10^{-34} \text{ kg m}^2/\text{s}$)

The larger the value for Δx is, the (larger/smaller) _____ the value for Δp is

- In other words, a **larger value for Δx** means we are **more uncertain about the position** of the particle and **less uncertain about the momentum** of the particle!

Max Born added to this and said that we can interpret this equation to mean that we cannot say exactly where electrons are but we can say where they are likely to be.

Example: Wave Particle Duality and the Heisenberg Uncertainty Principle

A beam of electrons is accelerated towards a thin metal foil and diffracts to give the exact same pattern as a blue laser with a wavelength of 480nm. The mass of an electron is 9.11×10^{-31} kg.

a) What speed were the electrons accelerated to?

b) If we measure the speed to a precision of $\pm 1\%$ what is the minimum uncertainty in the position of the electron?

3.10.3

Practice: Uncertainty of Position

A circus performer is shot out of a cannon and leaves the cannon with a velocity of 6.2 ± 0.1 m/s. If the performer weighs 82 kg, what is the minimum uncertainty in their position?



6.2 m

6.43×10^{-36} m

8.2 m

2.46×10^{-34} m

3.11 Schrodinger's Equation

3.11.1

Schrödinger Equation

$$\hat{\mathbf{H}}\psi = \mathbf{E}\psi$$

$\hat{\mathbf{H}}$: Hamiltonian

Ψ : Wavefunction

E: Total Energy

- For all but the simplest systems this equation cannot be solved analytically (exactly) but very good approximations can be obtained by computational methods.
- The wavefunction of a system, Ψ , can be used with a variety of operators to predict properties
- While the wavefunction does not have much of a “physical meaning”, the square of the wavefunction Ψ^2 gives the probability distribution of electron density.

Ψ : Wavefunction
 Ψ^2 : Probability Distribution

3.12 Polar Coordinates

3.12.1

Consequences of Moving to 3 Dimensions and Spherical Polar Coordinates

- When we move from 1-D to 3-D, we increase the complexity of the kinetic energy equations we discussed in the previous sections as well as the wavefunctions.

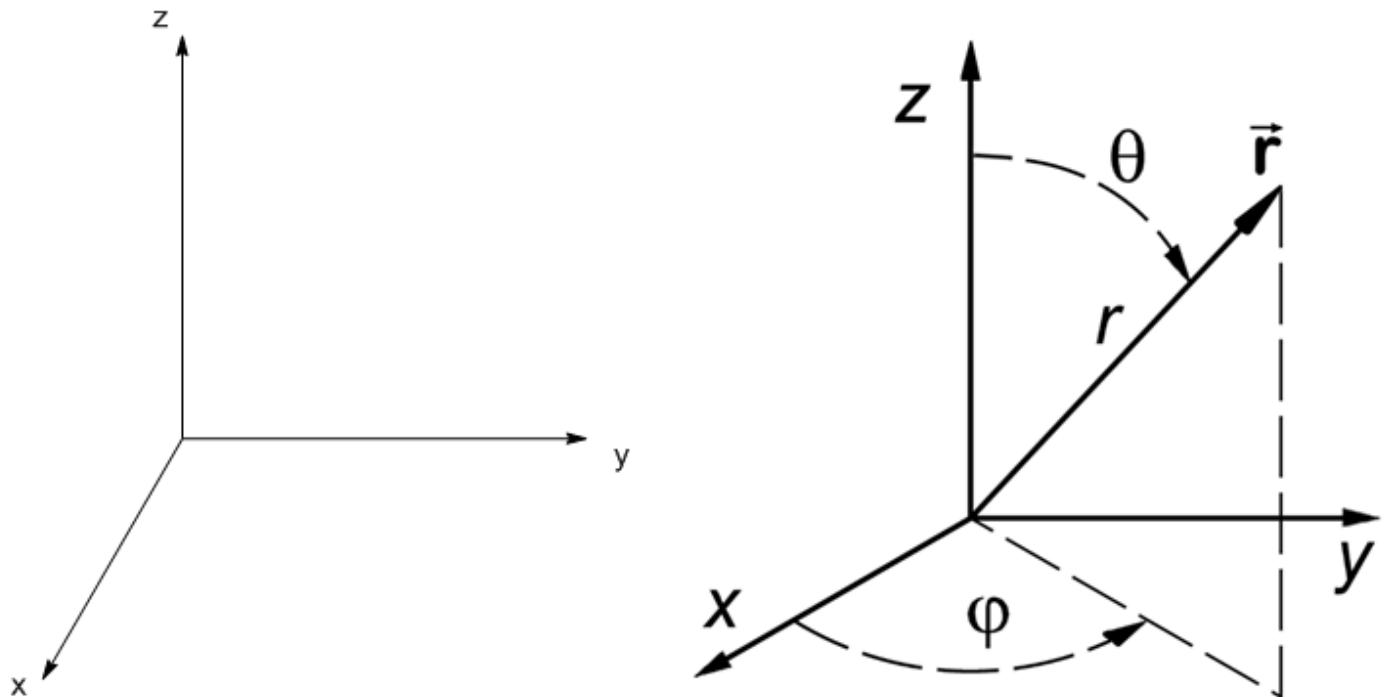
$$\mathbf{E}_k = \frac{\hbar^2}{8m} \left\{ \frac{n_x^2}{L_x^2} + \frac{n_y^2}{L_y^2} + \frac{n_z^2}{L_z^2} \right\}$$

- Schrödinger Equation, shown below, is what must be solved to obtain the wavefunction. This equation can only be solved exactly for one electron species and any multielectron species must be approximated. You are NOT required to know this equation.

$$-\frac{\hbar^2}{8\pi^2 m_e} \left\{ \frac{\delta^2 \psi}{\delta x^2} + \frac{\delta^2 \psi}{\delta y^2} + \frac{\delta^2 \psi}{\delta z^2} \right\} - \frac{Ze^2}{r} \psi = E\psi$$

- The wavefunction can be described as a function of (x,y,z), $\psi_{n,l,m_l}(x, y, z)$

- However, it is more convenient to change to spherical polar coordinates. These coordinates are a different set of variables that describe the same space as (x,y, z).



- Expressing the wavefunction in terms of spherical polar coordinates is advantageous because it allows the wavefunction to be split into a radial and angular component.

$$\psi_{n,l,m_l}(r, \theta, \psi) = R_{n,l}(r) \times Y_{l,m_l}(\theta, \psi)$$

3.13 The Quantum Model & Orbitals

3.13.1

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: $n = (1, 2, 3, 4\dots)$

- As n increases, **energy and size** of shell _____



WIZE TIP

n is also called the "principle quantum number"

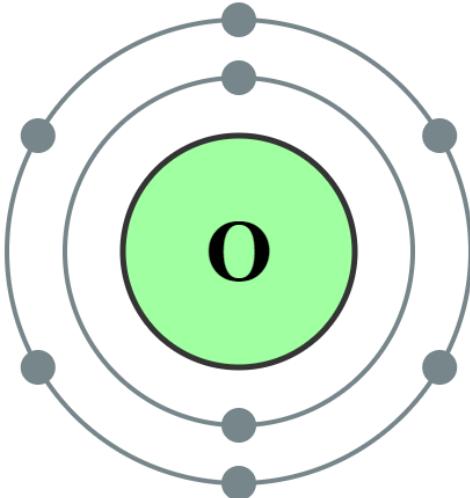


Photo by Greg Robson / CC BY

l

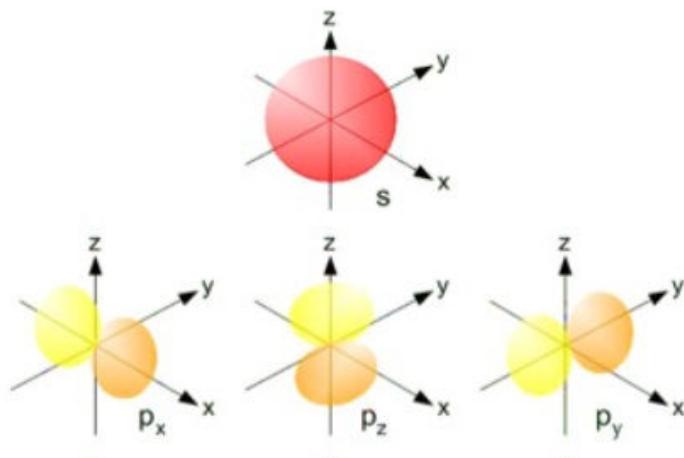
- Can be any **non-negative integer up to n-1**

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: $m_l = (-l, \dots, 0, \dots, l)$

- This quantum # **designates a specific orbital within a given shell**

Example: if $n=2$ and $l=1$, we are looking at $2p$.

- m_l 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 $-1/2$**

3.13.2

Quantum Numbers Summary

Principle Quantum Number (n)	Angular Momentum Quantum Number (ℓ)	Magnetic Quantum Number (m_ℓ)	Electron-Spin Quantum Number (m_s)	
$n = 1, 2, 3, \dots$	$\ell = 0, 1, 2, 3, \dots (n - 1)$	$m_\ell = \text{integer values of } \ell \text{ to } -\ell$	$m_s = \pm \frac{1}{2}$	
$n \uparrow \rightarrow E \uparrow$	$\ell = 0 \rightarrow s$ $\ell = 1 \rightarrow p$ $\ell = 2 \rightarrow d$	$\ell = 3 \rightarrow f$ $2\ell + 1$ orbitals	Only 2 electrons per orbital, ALWAYS	
n shell	ℓ subshell	Orbital	m_ℓ	
1	0	1s	0	1
2	0	2s	0	4
	1	2p	+1, 0, -1	
3	0	3s	0	
	1	3p	+1, 0, -1	9
	2	3d	+2, +1, 0, -1, -2	
4	0	4s	0	
	1	4p	+1, 0, -1	16
	2	4d	+2, +1, 0, -1, -2	
	3	4f	+3, +2, +1, 0, -1, -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) $(2, 0, 0, -\frac{1}{2})$

(ii) $(4, 3, 0, +\frac{1}{2})$

(iii) $(5, 1, 1, -\frac{1}{2})$

3.13.6

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, $-1/2$)

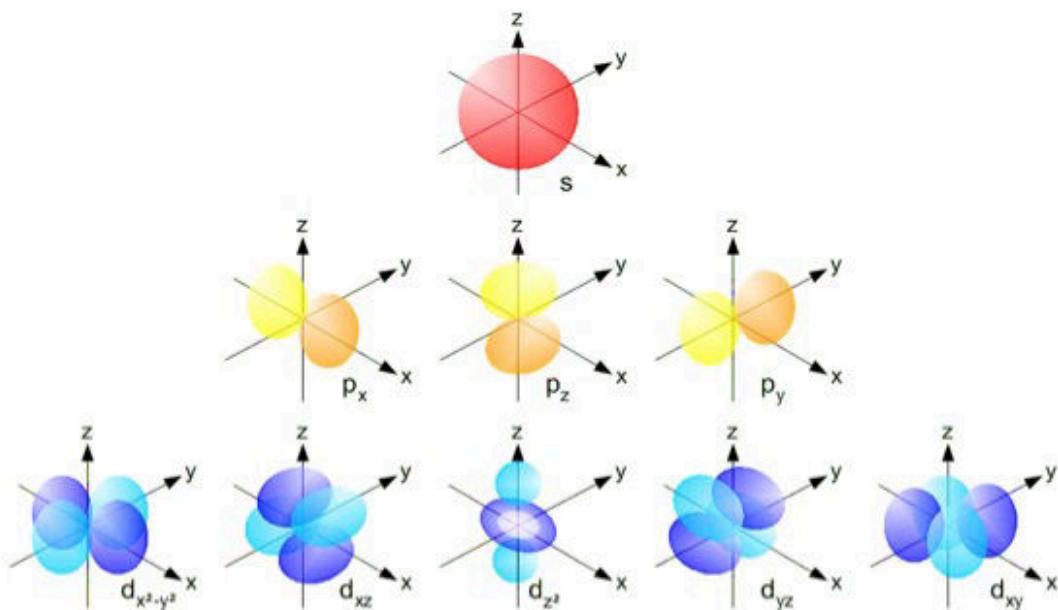
e) (3, 2, 3, $-1/2$)

f) (3, 1, 0, $+1/2$)

3.14 Shapes of Atomic Orbitals

3.14.1

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**)

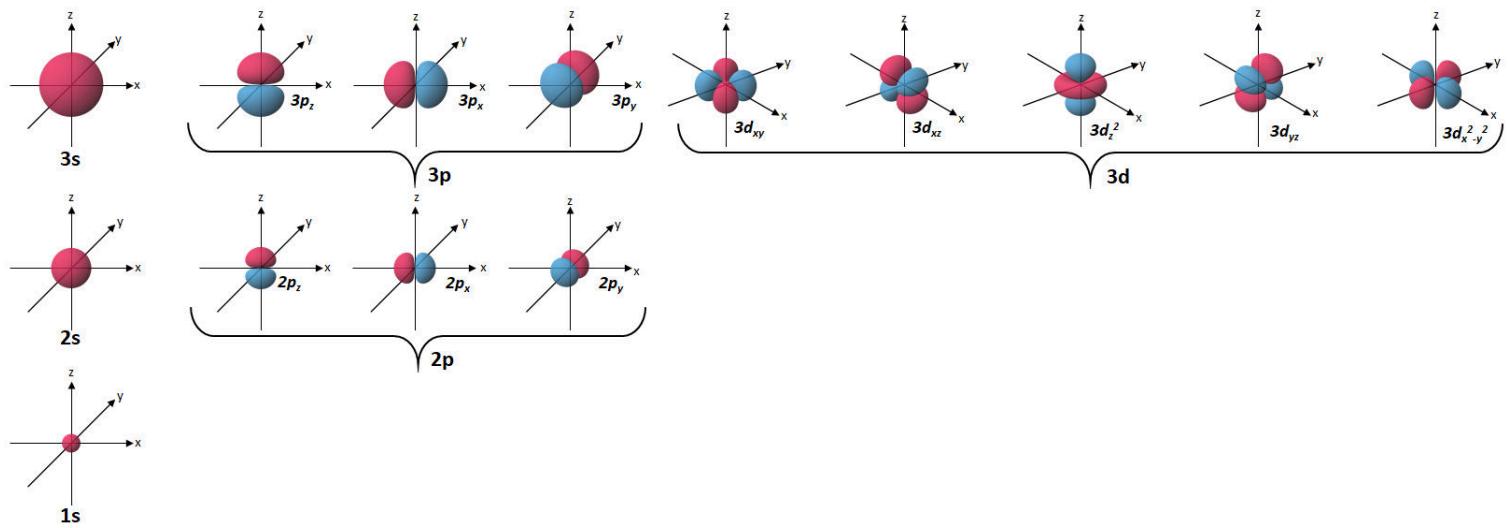


- There are **five d orbitals** in a subshell. The shapes of the d orbitals are more complicated than their s and p counterparts.
 - Three orbitals look like 3D cloverleaves, each lying in a plane with the lobes pointed between the axes (d_{xy} , d_{xz} and d_{yz}).
 - A fourth orbital is also a cloverleaf, 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

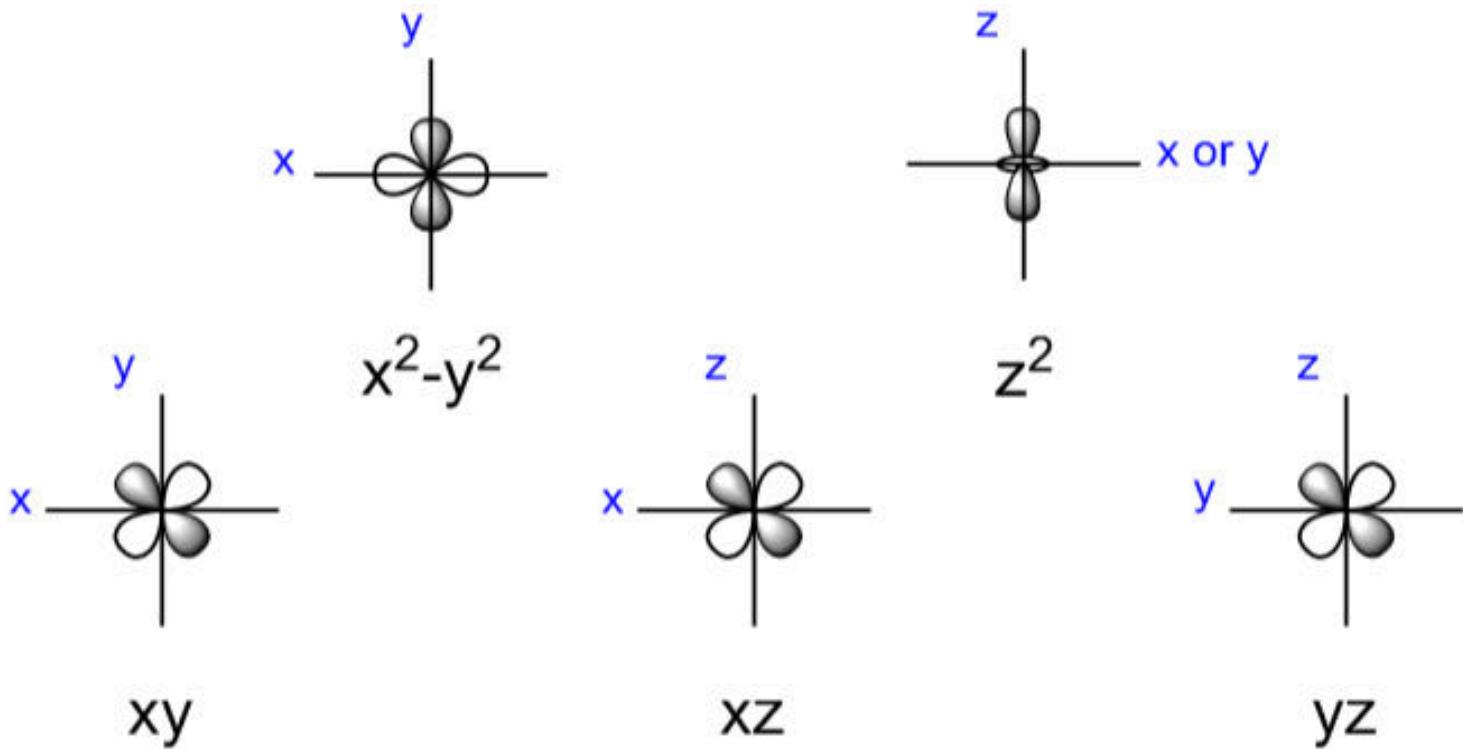
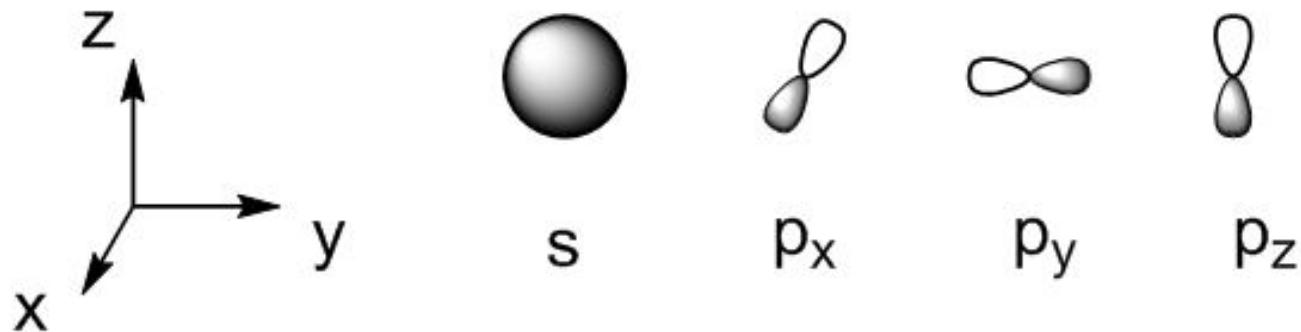
- As the **energy level increases**, so does the **number of orbitals**
Example: when n=1 (lower energy level), we see 1 orbital, whereas when n=2 (higher energy level), there are 4 possible orbitals
- As the **number of orbitals increases**, so does their **complexity**
Example: d orbitals are more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



3.14.2

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!

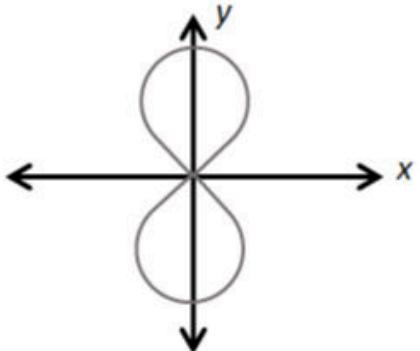
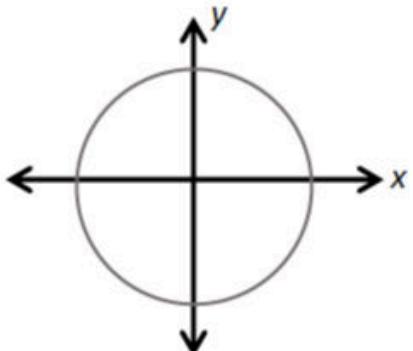
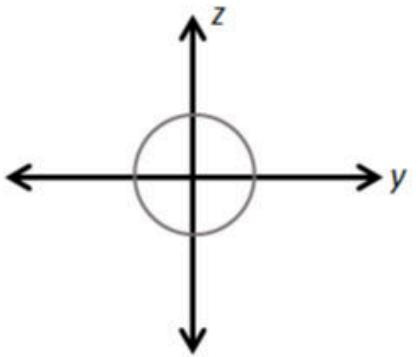
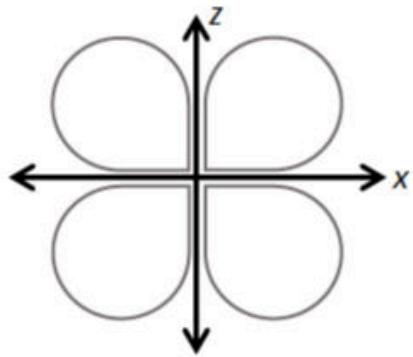
Also, notice how every other lobe is colored in gray. This gray and white coloring is indicating a "positive" lobe and a "negative" lobe. If you were asked to draw one of these orbitals on an exam, you would be expected to color in the lobes as well (or indicate which are + and which are -).

- 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 m_l is not associated with a particular orbital (ex. d_{xy} vs d_{xz}). Which means given a set of quantum numbers (n, l, m_l) one cannot determine which direction the orbital will be pointing

3.14.3

Example: Orbitals

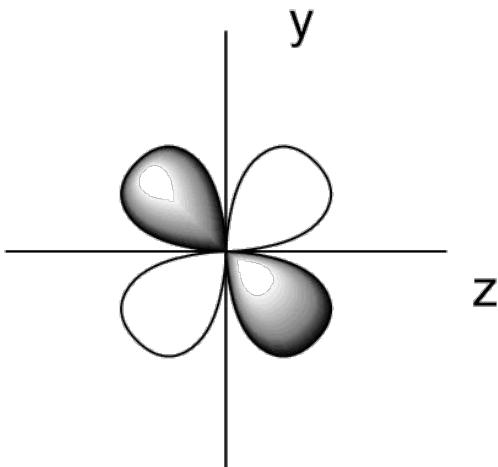
The sketches below show possible orbitals for the electron in a hydrogen atom. Which orbital would have the *lowest* energy?

- a.
- 
- b.
- 
- c.
- 
- d.
- 

3.14.4

Practice: Identify the Orbital

Identify the orbital shown below,



a) $2p_y$

b) $3d_{yz}$

c) $3p_x$

d) $3d_{xy}$

e) $3d_{x^2-y^2}$

3.15 Radial and Angulor Nodes

3.15.1

Angular and Radial Nodes

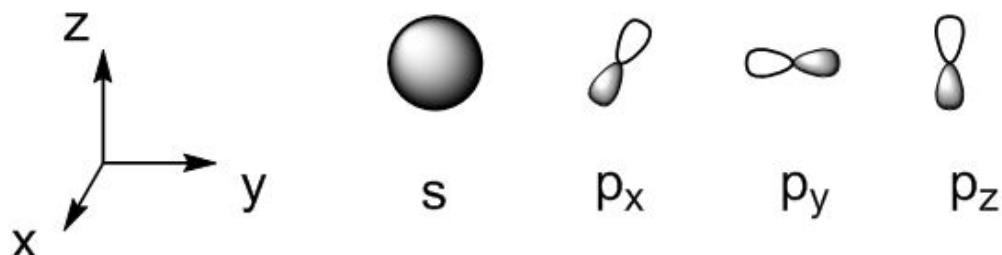
$$\text{Total number of nodes} = n - 1$$

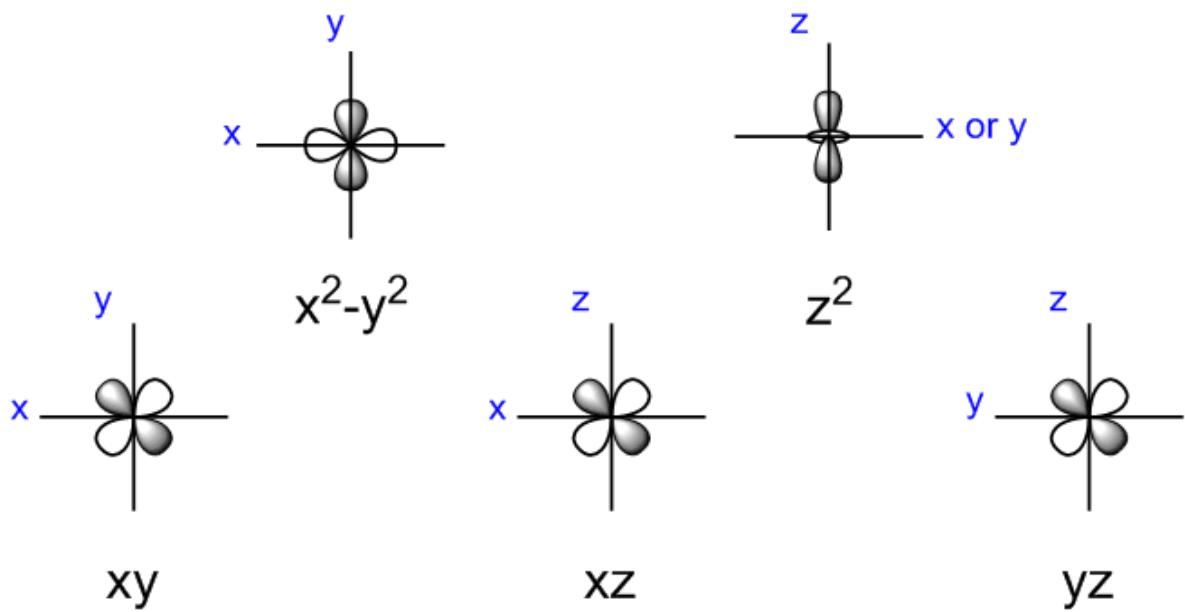
$$\text{Number of angular nodes} = l$$

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

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

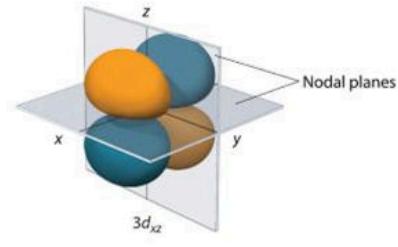
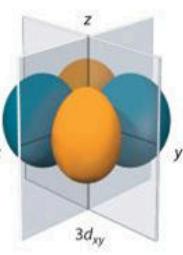
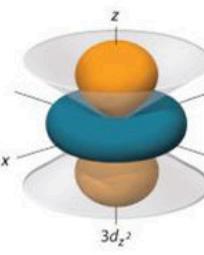
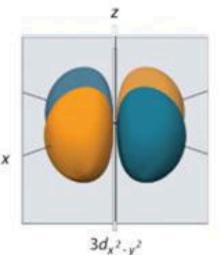
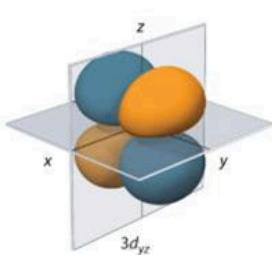
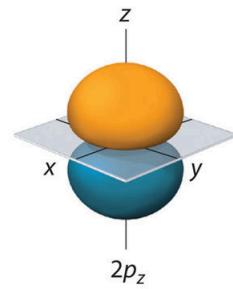
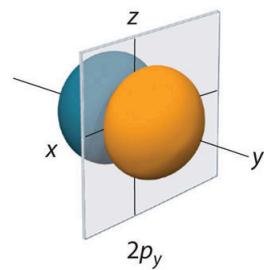
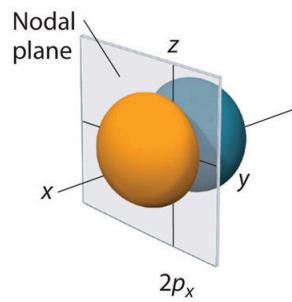
This time let's draw in the angular nodes for the 2p and 3d orbitals:





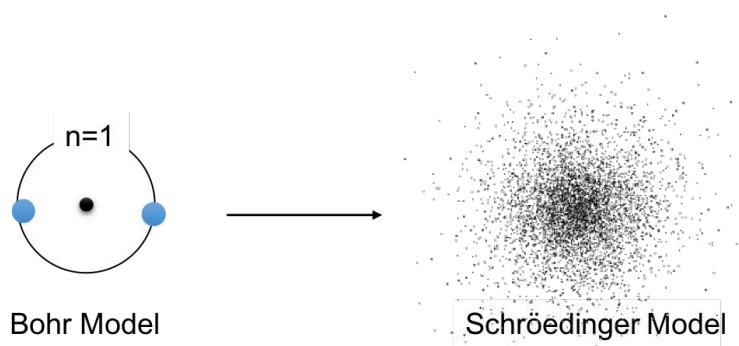
Another View of the Nodes

- **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...

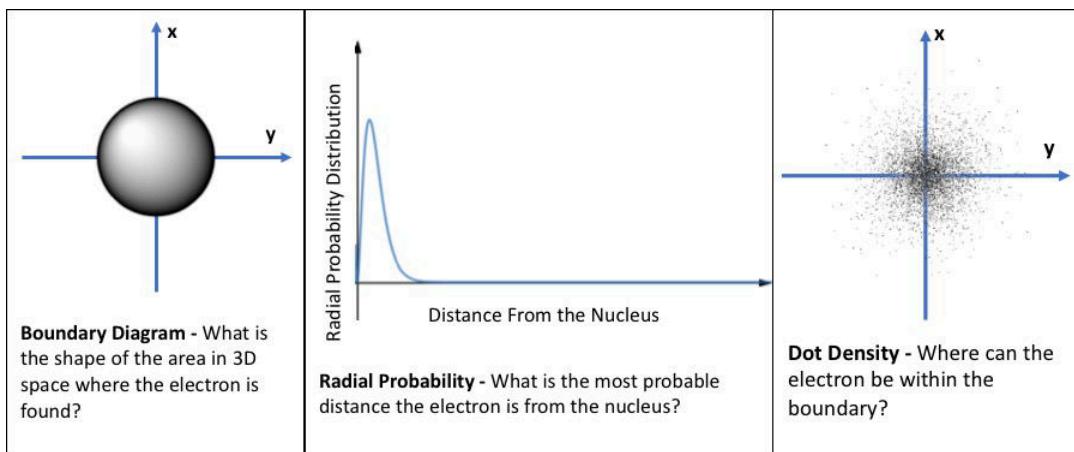


Wave Properties of Electrons

- 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



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**.

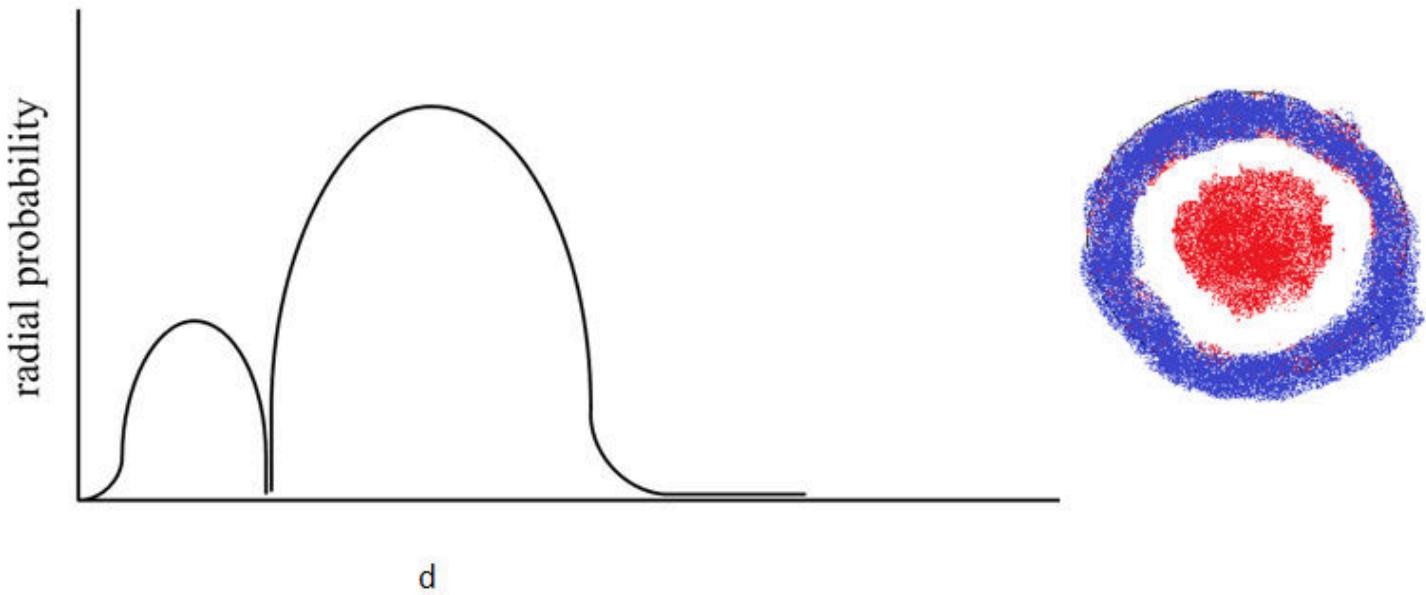
The only thing you will have need to “memorize” to do these problems is how to **draw each orbital**.

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!**

3.15.4

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

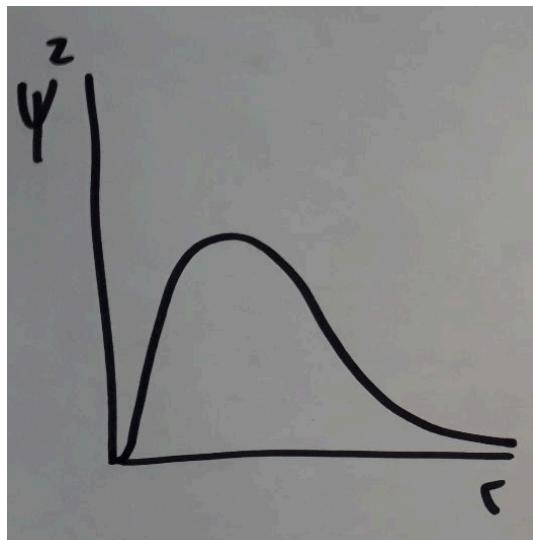
3.15.5

Example: Sketching Radial Probability

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

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?

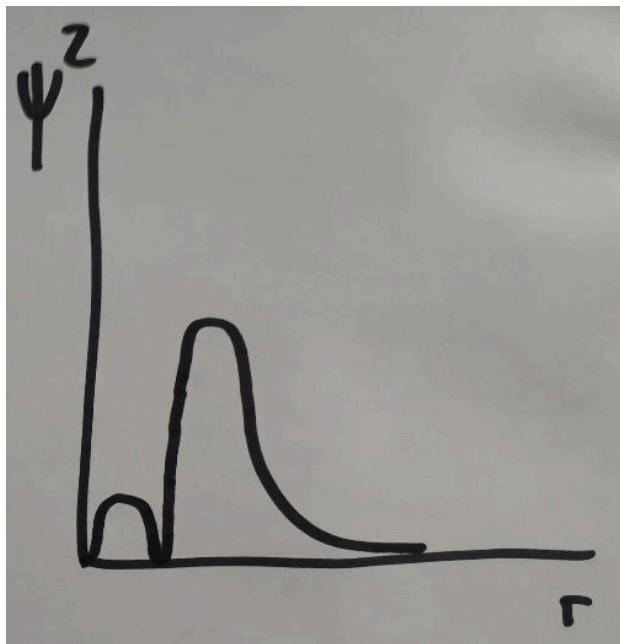


- a) $3d_{z^2}$
- b) $1s$
- c) $4p_x$
- d) $3s$
- e) $2p_y$

3.15.7

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

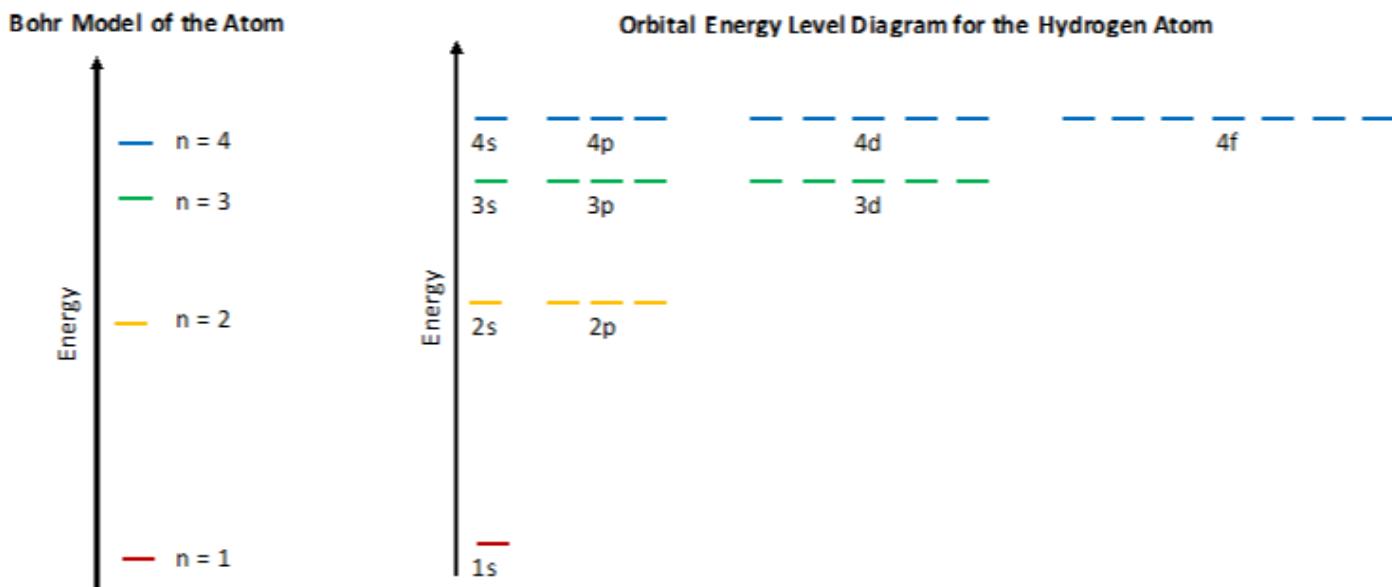
3.16 Orbital Filling Diagrams

3.16.1

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**.

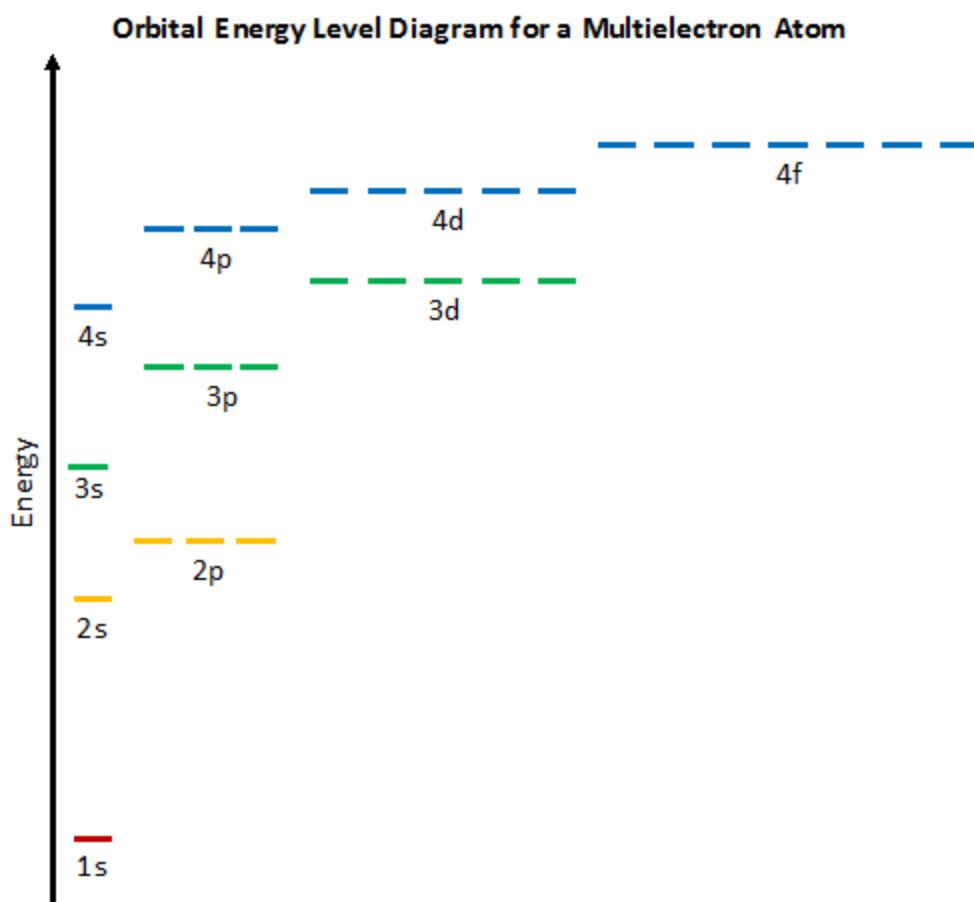


i WIZE TIP

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

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 _____

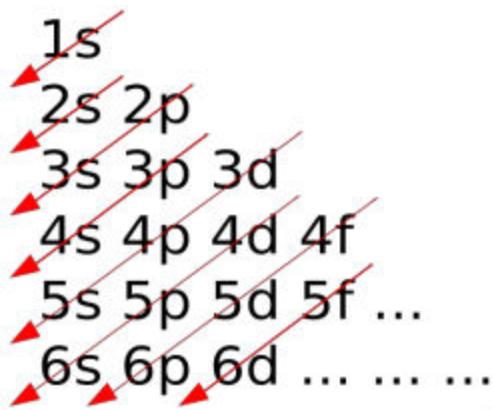
Rules for Orbital Filling

1) Aufbau Principle

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

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

				4f				
				5f				



! WATCH OUT!

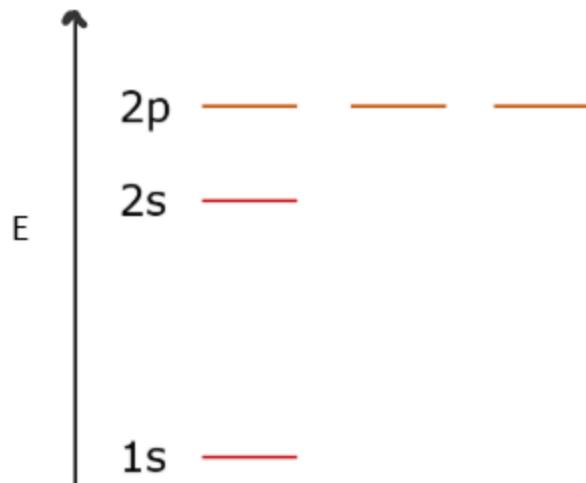
Memorize the order of orbital filling (or just familiarize yourself enough with the periodic table to know the order!): 1s, 2s, 2p, 3s, 3p, 3d, 4p...

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

3.16.3

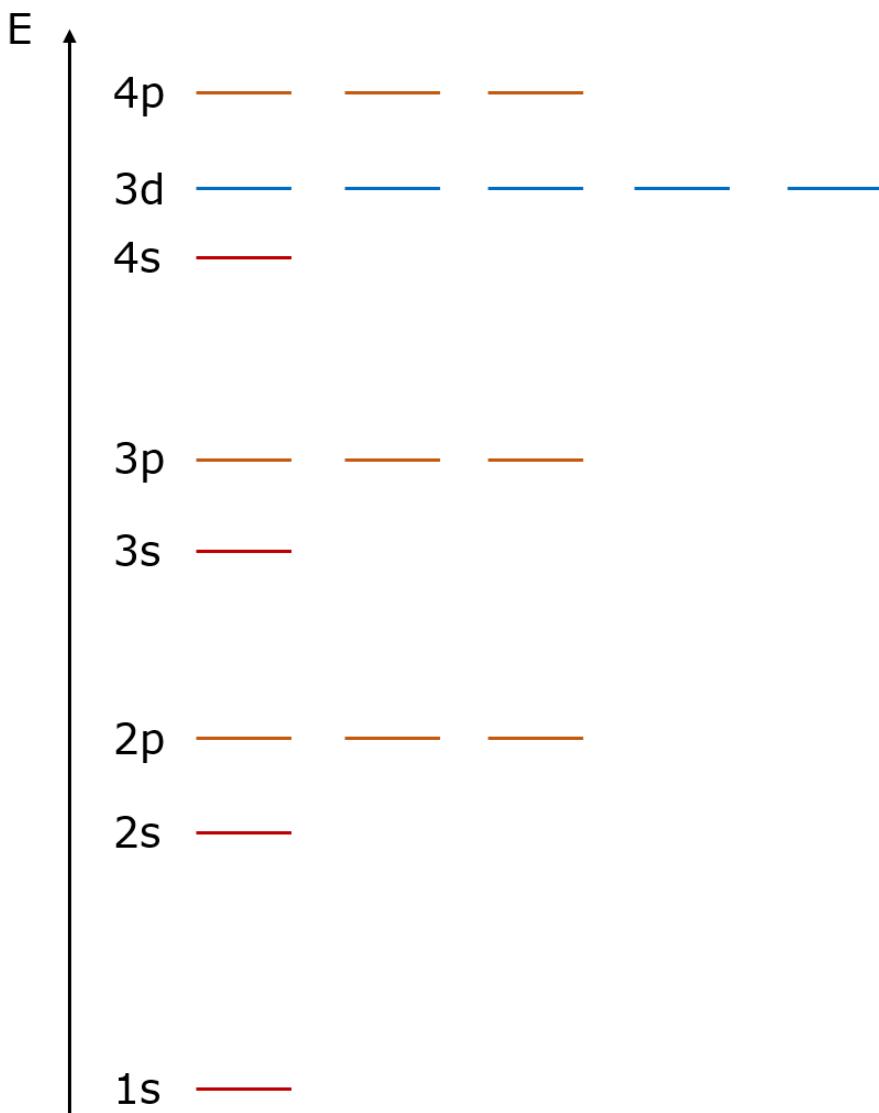
Example: Orbital Filling Diagrams

Draw the orbital diagram for oxygen.

3.16.4

Example: Orbital Filling Diagrams

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

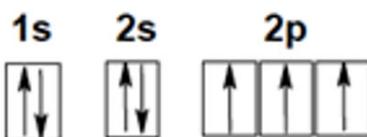


3.16.5

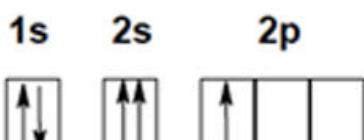
Practice: Orbital Filling Diagrams

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

A)



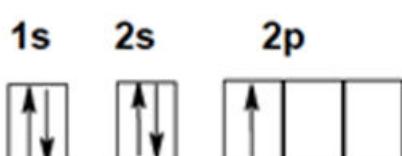
B)



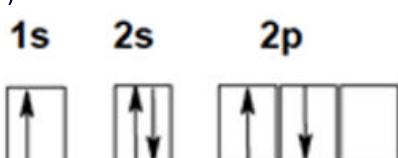
C)



D)



E)



3.16.6

Practice: Number of Orbitals

How many different atomic orbitals exist where $n = 3$?

3

6

9

12, which

18

3.17 Electronic Configuration for Atoms

3.17.1

Electron Configurations of Atoms

Group →	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
↓ Period																		
1	1 H																2 He	
2	3 Li	4 Be															10 Ne	
3	11 Na	12 Mg															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			

- I_nable blocks
- describe the n number for each block

Simple Electron Configurations

Li:

C:

Ne:

Group →	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
↓ Period																		
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:

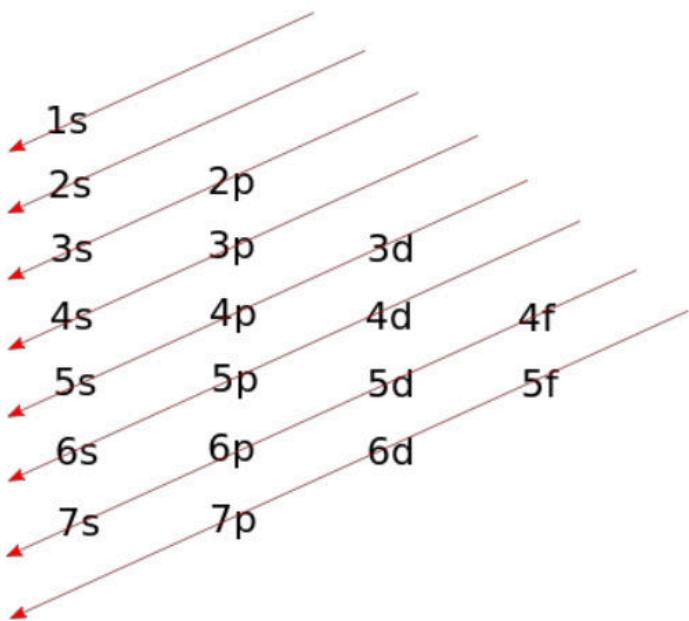
Fe:

Ge:

3.17.2

Electron Configurations Cheatsheet

1 H Hydrogen 1.008	2 He Helium 4.003
←1s→	
←2s→	
←3s→	
←4s→	3d→
←5s→	4d→
←6s→ *	5d→
←7s→ **	6d→
* Lanthanide series	
** Actinide series	
5 B Boron 10.81	6 C Carbon 12.011
7 N Nitrogen 14.012	8 O Oxygen 15.999
9 F Fluorine 18.998	10 Ne Neon 20.180
13 Al Aluminum 26.982	14 Si Silicon 28.085
15 P Phosphorus 30.974	16 S Sulfur 32.06
17 Cl Chlorine 35.45	18 Ar Argon 39.948
31 Ga Gallium 69.723	32 Ge Germanium 72.630
33 As Arsenic 74.922	34 Se Selenium 78.97
35 Br Bromine 79.904	36 Kr Krypton 83.798
51 Sb Antimony 122.80	52 Te Tellurium 127.60
53 I Iodine 126.904	54 Xe Xenon 131.293
61 Pb Lead 204.38	62 Bi Bismuth 208.980
63 Po Polonium (209)	64 At Astatine (210)
65 Rn Rhenium (223)	66 Mc Moscovium (223)
67 Ts Tennessine (223)	68 Lv Livermorium (223)
69 Yb Ytterbium 173.045	70 Lu Lutetium 174.924
57 Ce Cerium 138.905	58 Pr Praseodymium 140.918
59 Nd Neodymium 140.900	60 Pm Promethium 144.924
61 Sm Samarium 150.58	62 Eu Europium 151.964
63 Gd Gadolinium 157.95	64 Tb Terbium 158.925
65 Dy Dysprosium 162.500	66 Ho Holmium 164.900
67 Er Erbium 167.259	68 Tm Thulium 168.934
69 Yb Ytterbium 173.045	70 Lu Lutetium 174.924
89 Fr Francium (223)	90 Ra Radium (226)
91 Pa Protactinium 231.008	92 U Uranium 238.029
93 Np Neptunium (237)	94 Pu Plutonium (244)
95 Am Americium (243)	96 Cm Curium (247)
97 Bk Berkelium (247)	98 Cf Californium (251)
99 Es Einsteinium (257)	100 Fm Fermium (257)
101 Md Mendelevium (256)	102 No Neptunium (259)



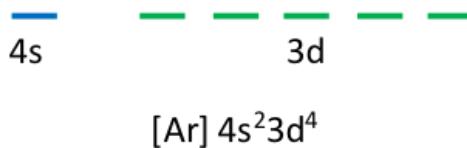
Electron Configuration Exceptions

Transition metals have their valence electrons in the d subshell.

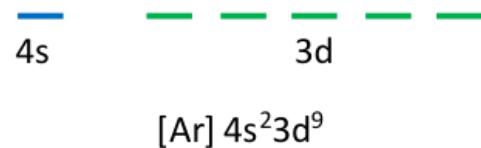
It is a lot **more favourable** (and stable) to have **d shell either half full or totally full instead of just partially full.**

Cr:

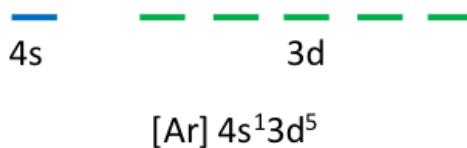
Expected:



Cu:



Seen:



i **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!)

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:

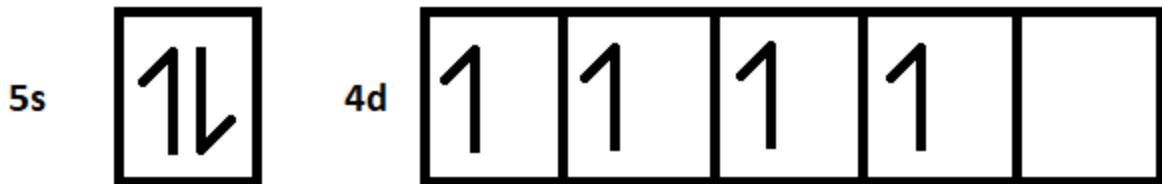
Group →	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
↓ Period																		
The Periodic Table of the Elements																		
1	1 H															2 He		
2	3 Li	4 Be															10 Ne	
3	11 Na	12 Mg															18 Ar	
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 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																		
Actinides																		
	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			
	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

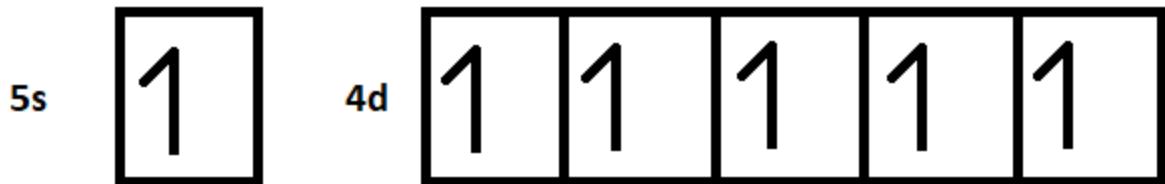
What is the condensed electronic configuration for Mo?

	Group → 1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
↓ Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
The Periodic Table of the Elements																		
1	1 H																2 He	
2	3 Li	4 Be															10 Ne	
3	11 Na	12 Mg															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																		
Actinides																		
	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			
	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			

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]5s¹4d⁵**.

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



3.17.7

Practice: Electron Configuration

The ground state electronic configuration for At is

	Group → 1 ↓ Period	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) $[Xe]6s^24f^{14}5d^{10}6p^5$

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

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

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

3.17.8 **Practice**

Practice: Impossible Electron Configuration

Which of the following electronic configurations is not possible?

Group → 1 ↓ Period	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
The Periodic Table of the Elements																	
1	1 H																2 He
2	3 Li	4 Be															10 Ne
3	11 Na	12 Mg															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
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
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	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	118 Og
Lanthanides																	
Actinides																	
	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		
	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		

[Ar]4s²4d⁶

1s²2s²2p⁶3s²3p⁶4s¹

[Kr]5s¹

1s²2s²2p⁶3s²3p⁶4s²3d⁸

3.18 Electron Configurations for Ions

3.18.1

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: 2s has a higher energy than 1s

In order of **increasing energy for the subshells**, we write:

s < p < d < f

Example: 2p has a higher energy than 2s

F: $1s^2 2s^2 2p^5$



[He] $2s^2 2p^5$

F: $1s^2 2s^2 2p^6$



[Ne]

What is the electron configuration for F^+ ?

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!

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!

Fe: [Ar] 4s²3d⁶



Fe¹⁺: [Ar] 4s¹3d⁶



Fe²⁺: [Ar] 3d⁶



Fe³⁺: [Ar] 3d⁵



i 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!

3.18.3

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?

Group →	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
↓ Period																		
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			

i) N⁺ and C

ii) C and B⁺

iii) Cl^- and Ar

3.18.4

Example: Writing Electron Configurations of Ions

Write the electronic configuration for the following species:

Group → 1 ↓ Period	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
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
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
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
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+}

3.18.5

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

3.19 Diamagnetic vs Paramagnetic Electron Configurations

3.19.1

Diamagnetic vs Paramagnetic

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

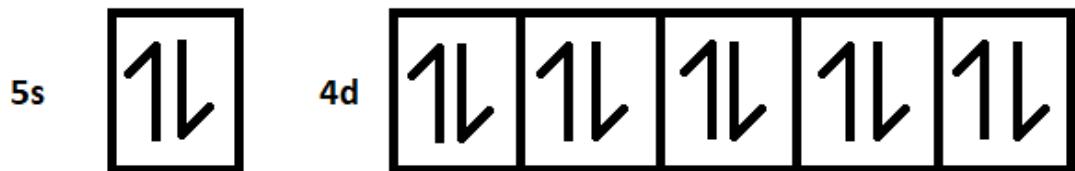
Diamagnetic - all electrons are **paired**

Paramagnetic - all electrons are **not paired**

Are the following electron configurations diamagnetic or paramagnetic?

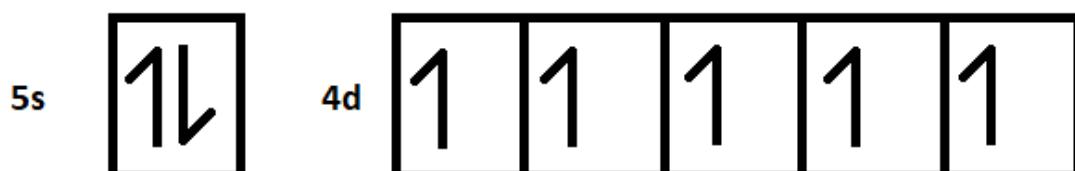
Cd

[Kr] 5s² 4d¹⁰



Tc

[Kr] 5s² 4d⁵



Group →	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
↓ Period																		
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.

3.19.2

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+}

3.19.3

Practice: Diamagnetic

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

Group →	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
↓ Period																		
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																		
Actinides																		
	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			
	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			

Cr

Co

Ru

None of the above

[View Solutions on Wizeprep.com](https://www.wizeprep.com)

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

3.20 Effective Nuclear Charge (Zeff)

3.20.1

Effective Nuclear Charge (Zeff)

To understand effective nuclear charge (Zeff), let's consider a concert:



Photo by The Come Up Show / CC BY

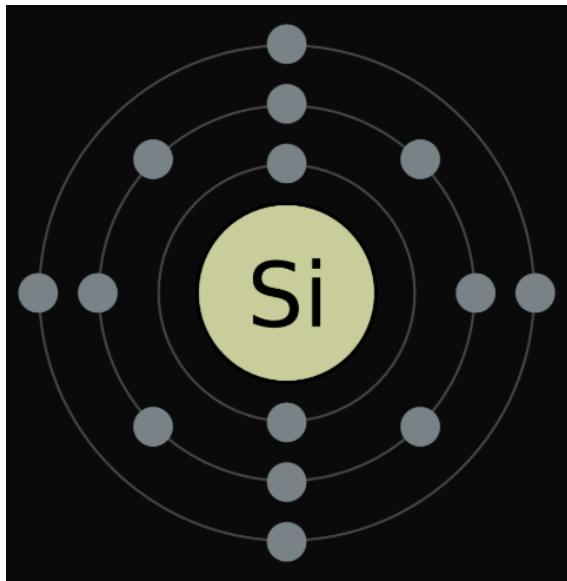
- People closer to the stage at a concert are going to be more into it. The music is louder and they really get a close connection with the artist. People in the highest rows in the stadium get less of that connection with the artist performing.
- This helps to explains the concept of nuclear shielding. **Inner electrons** (or people sitting in the front rows in our example) **shield the outer electrons** (or people sitting in the highest rows) from the **attractive force of the nucleus** (or artist in our example)

Effective Nuclear Charge

This is the **nuclear charge that is "felt"** by a valence electron.

Core Electrons Vs Valence Electrons

Label the **core** and **valence electrons** in the diagram below:



- Valence electrons are attracted to the positively charged nucleus BUT valence electrons are repelled by the core electrons
- Note that electrons in the same shell "feel" the same attraction to the nucleus (since they are the same distance from the nucleus, just like how the people in the same row would feel the same connection to the artist)

$$Z_{eff} = Z - S$$

Z_{eff} is the **effective nuclear charge**
Z is the **atomic number** (# of protons)
S is the # of **shielding electrons**

Are the shielding electrons core electrons or valence electrons? _____

What would be the effective nuclear charge for Lithium?

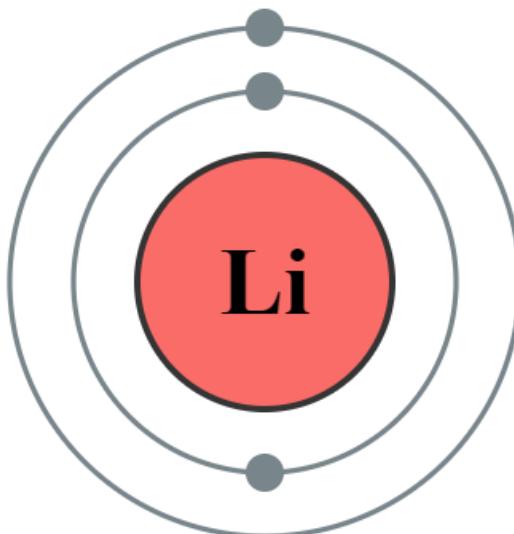


Photo by Greg Robson / CC BY

The Periodic Trend

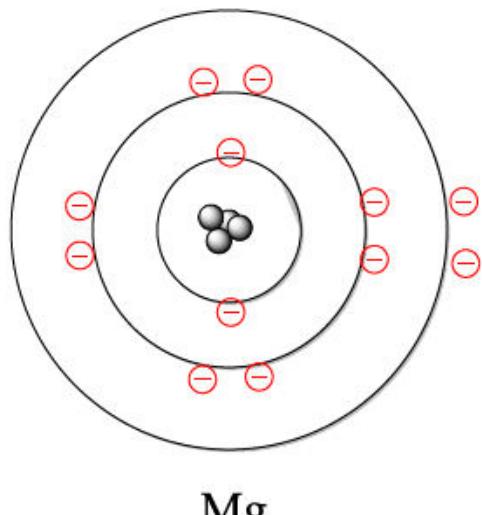
- As we **move to the right** across the periodic table, the # of core electrons stays the same but the **# of protons increases**. Therefore, Z_{eff} **increases**.
 - With protons, can pull the electrons in closer → Z_{eff}
 - Effects of shielding are less with more protons
- As we **move down a group** of the periodic table, the **# of core shells increases** and the valence electrons get further from the nucleus. This increased distance from the nucleus leads to a **smaller Z_{eff}** .
 - a group → shells
 - shells → shielding
 - shielding → Z_{eff}

Hydrogen 1 H	Lithium 3 Li	Boron 5 B	Carbon 6 C	Nitrogen 7 N	Oxygen 8 O	Sulfur 16 S	Neon 10 Ne
Helium 2 He	Be	Silicon 14 Si	Phosphorus 15 P	Sulfur 16 S	Chlorine 17 Cl	Ar	
Li	Magnesium 12 Mg	Aluminum 13 Al	Silicon 14 Si	Phosphorus 15 P	Chlorine 17 Cl	Sulfur 16 S	
Na	Calcium 20 Ca	Scandium 21 Sc	Titanium 22 Ti	Vanadium 23 V	Chromium 24 Cr	Manganese 25 Mn	Ferrum 26 Fe
K	Sodium 19 K	Scandium 21 Sc	Titanium 22 Ti	Vanadium 23 V	Chromium 24 Cr	Manganese 25 Mn	Iron 26 Fe
Rb	Rubidium 37 Rb	Yttrium 39 Y	Zirconium 40 Zr	Nickel 41 Ni	Chromium 24 Cr	Manganese 25 Mn	Iron 26 Fe
Cs	Ce 55 Cs	La 57 La	Hafnium 72 Hf	Pt 78 Pt	Gold 79 Au	Hg 80 Hg	Tl 81 Tl
Fr	Ra 87 Fr	Lanthanides 88-103 Lr	Rutherfordium 104 Rf	Rhenium 106 Re	Hassium 108 Hs	Mendelevium 109 Mts	Ununtrium 110 Uun
						Ununpentium 111 Uuu	Ununhexium 112 Uub
						Ununseptium 113 Uus	Ununoctium 114 Uuo

3.20.2

Example: Estimating Z_{eff}

What is the Z_{eff} of an electron in the $n=3$ shell of magnesium (Mg)?



3.20.3 **Practice**

Practice: Estimating Z_{eff}

Calculate the approximate effective nuclear charge of Se.

Group →	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
↓ Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
	1 H																2 He	
1	3 Li	4 Be															9 F	10 Ne
2	11 Na	12 Mg															17 Cl	18 Ar
3	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
4	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
5	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
6	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
7	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		

6



8



28



34



3.21 Atomic Radius Trend

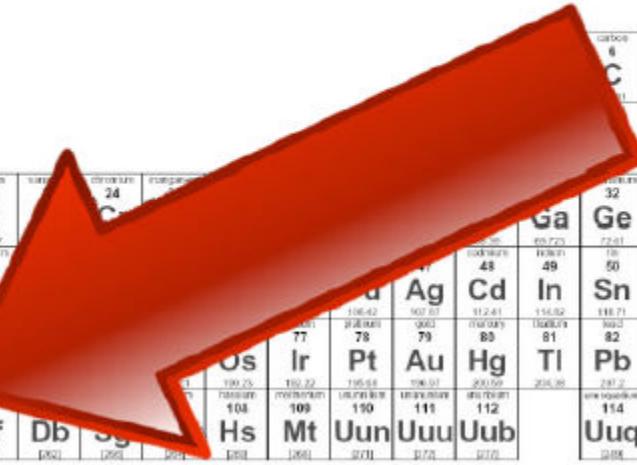
3.21.1

Atomic Radius

The atomic radius is the estimated **radius of an atom** (from the nucleus to the outermost valence electrons)

The Periodic Trend

- As we **move to the right** on the periodic table, the **Z_{eff} increases** and this “pulls” the electrons closer to the nucleus which **decreases the radii**.
- As we **move down a group** in the periodic table, the number of **electron shells increases** which makes the **atom radii larger**.
 - group → shells
 - shells → shielding and Z_{eff}
 - With a $Z_{\text{eff}} \rightarrow$ pull on outer electrons, leading to atomic radii



hydrogen	1	H	1.0070
lithium	3	Li	6.941
boron	4	Be	9.0122
nitrogen	5		
magnesium	11	Na	22.990
sulfur	12	Mg	24.305
phosphorus	15	K	39.098
chlorine	19	Ca	40.078
iodine	37	Rb	85.451
strontium	38	Sr	87.62
tin	55	Cs	118.71
barium	56	Ba	137.33
francium	87	Fr	223.02
radium	88	Ra	226.02
lanthanides	89-102	*	*
actinides	103-105	Lr	174.97
curium	104	Rf	174.97
neptunium	105	Db	174.97
curium	106	Sg	174.97
hassium	108	Hs	174.97
meitnerium	109	Mt	174.97
ununbium	111	Uun	174.97
ununquadium	112	Uub	174.97
ununpentium	114	Uup	174.97
helium	2	He	4.0030
carbon	6	C	12.011
nitrogen	7	N	14.007
oxygen	8	O	15.999
fluorine	9	F	18.998
phosphorus	15	P	30.974
sulfur	16	S	32.065
chlorine	17	Cl	35.455
aristotele	18	Ar	36.048
silicon	14	Si	30.974
silicon	14	Ge	72.61
arsenic	33	As	74.72
antimony	34	Se	77.98
tellurium	35	Br	79.904
iodine	36	Kr	83.80
astatine	53	Xe	131.29
astatine	54	I	132.91
astatine	85	Po	136.91
astatine	86	At	139.91
astatine	121	Rn	162.91

Element with the **largest atomic radii** in the periodic table: _____

Element with the **smallest atomic radii** in the periodic table: _____

How do ions change this?

- **Anions:** _____ due to increased electron-electron repulsion
- **Cations:** _____ due to decreased electron-electron repulsion
 - How does this relate to Z_{eff}?

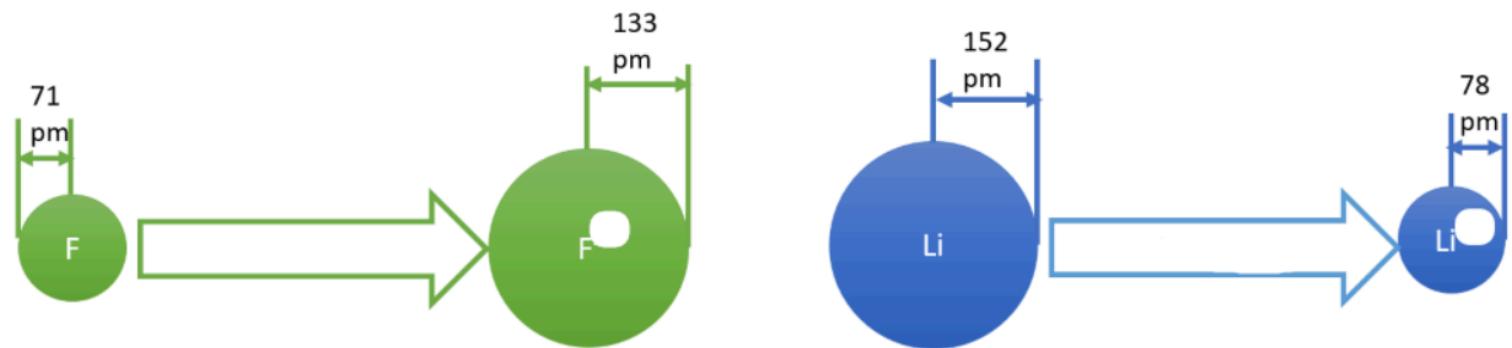
In general:

Ionic Radius: Three Scenarios

You could be asked to rank different atoms or ions according to the sizes of their ionic radii using the trends we discussed.

Here are three common scenarios you may encounter:

1. Same element different charge:



2. Different element same charge:

- Identical trend to atoms

Example: Li < Na so $\text{Li}^+ < \text{Na}^+$

3. Different element different charge

- Can only be assessed for **isoelectronic species** (same # of electrons)
- Compare the **proton to electron ratio**.
 - **More protons than electrons**, means stronger pull, **smaller radius**.
 - **Less protons than electrons**, means weaker pull, **larger radius**.

Example: Compare the radii for the following: O²⁻ F⁻ Ne Na⁺ Mg²⁺

Group →	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
↓ Period																		
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			

3.21.3

Example: Ranking Size of Atoms

Rank the following atoms in order of increasing atomic radius: Se, Cs, Br, Ga, F, As.

Group →	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
↓ Period																		
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		

3.21.4

Example: Ranking Size of Ions

Rank the following species in order of decreasing size. F^- , N^{3-} , Al^{3+}

Group →	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
↓ 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			

i WIZE TIP

When comparing atoms that are **isoelectronic** (i.e. they have the same number of electrons), remember **more protons in the nucleus** means a stronger pull of electrons toward the + charged nucleus, so the **smaller the atomic radius**

3.21.5

Practice: Determining the Largest Atom

Select the atom below which has the largest atomic radius?

Group →	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
↓ Period																		
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			

S

I

O

In

Rb

[View Solutions on Wizeprep.com](#)

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

3.22 Ionization Energy

3.22.1

Ionization Energy (IE)

The ionization energy is the amount of **energy required to remove the outermost electron** from an atom or ion in the gas phase

- If removing an electron do we get an anion or a cation? _____



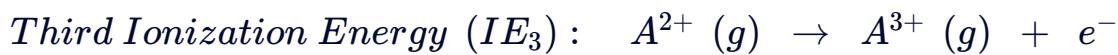
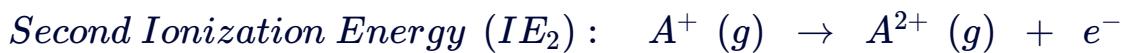
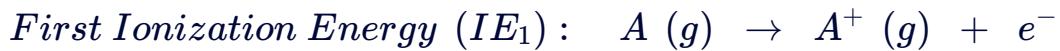
- ↑IE means that the atom/ion requires **more energy** for a valence **electron to be removed** (has a tight pull on electron)

The Periodic Trend

- As we **move to the right** across a period, **Z_{eff} increases** and the electrons are held more tightly. Therefore, it takes more energy to remove an electron and the **IE increases** as we move to the right.
- As we **move down a group** and the valence electrons are further away from the nucleus, they are held more weakly (**lower Z_{eff}**) and **IE decreases** (it becomes easier to remove the outermost electron)

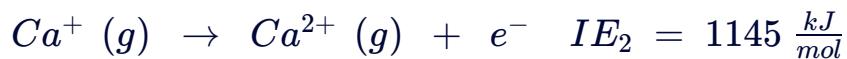
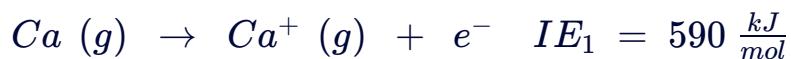
Comparing 1st, 2nd, and 3rd Ionization Energies

How does the first IE compare to the second IE?



1st ionization is always smaller than the 2nd because removing a negatively charged electron from a cation is more difficult.

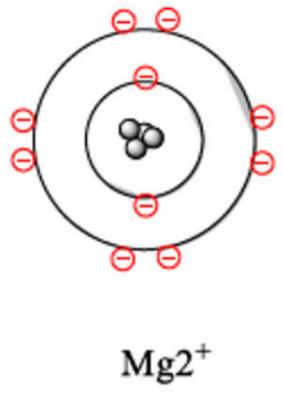
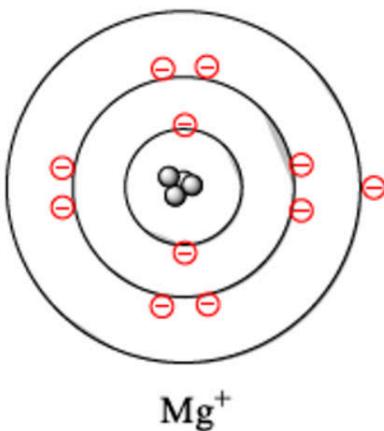
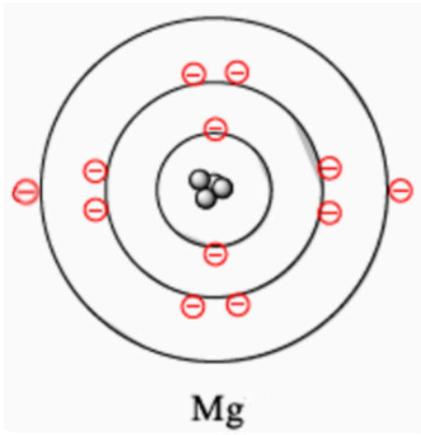
Example:



WIZE TIP

$IE_1 < IE_2 < IE_3 < IE_4$ etc

How does the IE_2 and IE_3 for magnesium compare?

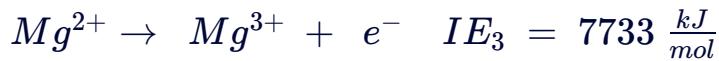
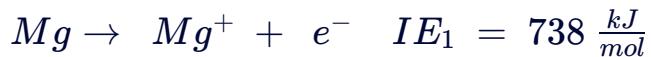


A **core electron is much more difficult to remove than a valence electron**, for example: the 3rd ionization energy of magnesium is much, much, higher than the 2nd.

i **WIZE TIP**

Although the IE increases each time another electron is removed, the increase is not linear and is usually related to the electron configuration.

Ionizing Mg²⁺ would be very difficult!



Needing so much energy to remove the 3rd electron indicates that when 2 were removed it was stable/unstable: _____ !

3.22.3

Practice: Highest 1st Ionization Energy

Which of the following species has the highest 1st ionization energy?

Group →	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
↓ Period																		
1	1 H																2 He	
2	3 Li	4 Be															10 Ne	
3	11 Na	12 Mg															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			

S

F

B

N

P

[View Solutions on Wizeprep.com](https://www.wizeprep.com)

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

Practice: Trends in Ionization Energy

Rank the following atoms in order of increasing first ionization energy (1 = smallest; 6 = largest): Fe, O, Ba, Si, V, Zr

Group →	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
↓ Period																		
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		

Fe	_____
O	_____
Ba	_____
Si	_____

V	
Zr	

3.22.5

Practice: Lowest Ionization Energy

Select the element with the lowest 3rd ionization energy.

Group →	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
↓ Period																		
1	1 H																2 He	
2	3 Li	4 Be															10 Ne	
3	11 Na	12 Mg															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

B) Mg

C) Cl

D) I

E) C

3.23 Electron Affinity Trend

3.23.1

Electron Affinity (EA)

Electron affinity is the **amount of energy involved with adding an electron** to an atom or ion.

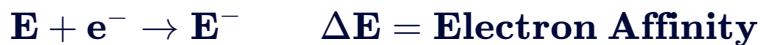
 **WIZE TIP**

Don't confuse ionization energy (IE) with electron affinity (EA)!

- **Ionization energy** is when we **remove** the outermost electron
- **Electron affinity** is when we **add** an electron

Memory tip: Electron affinity has an "a" in it for adding an electron

When adding an electron, do we get a cation or an anion? _____ !



- **If EA is negative** → energy is released/absorbed: _____
 - If this is the case, would the atom go to a higher/lower E state? _____
 - This means the atom has become more/less _____ stable as a result of adding an electron
 - Did this atom like or dislike having an electron added to it? _____
- **If EA is positive** → energy is released/absorbed: _____
 - If this is the case, would the atom go to a higher/lower E state? _____
 - This means the atom has become more/less stable _____ stable as a result of adding an electron
 - Did this atom like or dislike having an electron added to it? _____

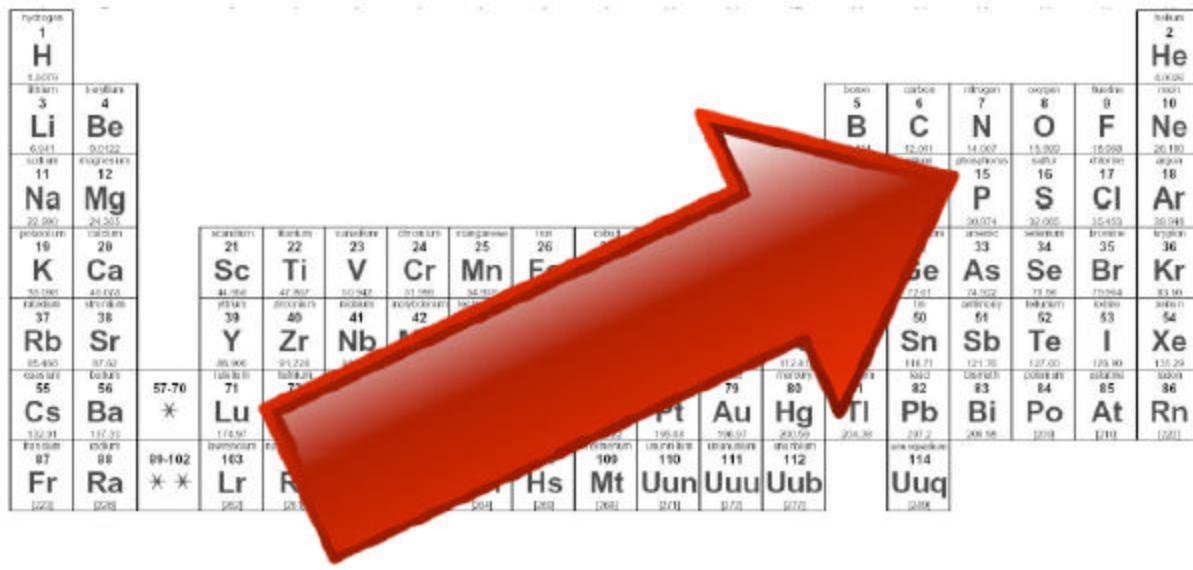
! WATCH OUT!

Use a **negative EA** as a reference point, unless the question states otherwise.

- This means that unless specified otherwise, when asked which element would have the largest electron affinity, assume they mean largest NEGATIVE electron affinity.
 - This would mean that the element really liked getting an electron and lost energy, resulting in a more stable form.

The Periodic Trend

In general, as Z_{eff} increases, the incoming electron experiences a greater electrostatic attraction and **stabilization** leading to a **greater (negative) electron affinity**



What about noble gases?

Do you think a noble gas would have a positive or negative electron affinity? _____

- Are noble gases stable? _____
- Would a noble gas want another electron added? _____
- If you tried to add an electron to a noble gas would that require energy or release energy?

 WATCH OUT!

This trend excludes noble gases. Noble gases have stable, completely filled shells. Adding electrons to noble gases will break the noble gas configuration.

3.23.2

Example: Increasing Electron Affinity

Rank the following atoms in increasing absolute energy of their first electron affinity: Br, Ag, Ba, Mo, Sb

Group →	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
↓ Period																		
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			

3.23.3

Practice: Positive Electron Affinity

For most elements, the first electron affinity is negative; however, a few elements have a positive first electron affinity. Which atom has the most positive first electron affinity?

F

B

Ne

C

O

Na

3.23.4

Practice: Highest Electron Affinity

Which of the following has the highest electron affinity?

N



N+



N-



Not enough information is provided to determine this.



3.24 Electronegativity

3.24.1

Electronegativity (EN)

Electronegativity is a **measure of electron pull**. In a bond, the **more electronegative** element will **pull electrons closer towards itself**

Example: Draw the bond between H and F and show the difference in electronegativity.

Group →	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
↓ Period	H 2.20	Li 0.98	Be 1.57															He
1																		
2													B 2.04	C 2.55	N 3.04	O 3.44	F 3.98	Ne
3													Al 1.61	Si 1.90	P 2.19	S 2.58	Cl 3.16	Ar
4	K 0.82	Ca 1.00	Sc 1.36	Ti 1.54	V 1.63	Cr 1.66	Mn 1.55	Fe 1.83	Co 1.88	Ni 1.91	Cu 1.90	Zn 1.65	Ga 1.81	Ge 2.01	As 2.18	Se 2.55	Br 2.96	Kr 3.00
5	Rb 0.82	Sr 0.95	Y 1.22	Zr 1.33	Nb 1.6	Mo 2.16	Tc 1.9	Ru 2.2	Rh 2.28	Pd 2.20	Ag 1.93	Cd 1.69	In 1.78	Sn 1.96	Sb 2.05	Te 2.1	I 2.66	Xe 2.60
6	Cs 0.79	Ba 0.89	*	Hf 1.3	Ta 1.5	W 2.36	Re 1.9	Os 2.2	Ir 2.20	Pt 2.28	Au 2.54	Hg 2.00	Tl 1.62	Pb 1.87	Bi 2.02	Po 2.0	At 2.2	Rn 2.2
7	Fr 0.7	Ra 0.9	**	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Cn	Uut	Fl	Uup	Lv	Uus	Uuo
* Lanthanoids		La 1.1	Ce 1.12	Pr 1.13	Nd 1.14	Pm 1.13	Sm 1.17	Eu 1.2	Gd 1.2	Tb 1.1	Dy 1.22	Ho 1.23	Er 1.24	Tm 1.25	Yb 1.1	Lu 1.27		
** Actinoids		Ac 1.1	Th 1.3	Pa 1.5	U 1.38	Np 1.36	Pu 1.28	Am 1.13	Cm 1.28	Bk 1.3	Cf 1.3	Es 1.3	Fm 1.3	Md 1.3	No 1.3	Lr 1.3		

Electronegativity Values

Photo by DMacks / CC BY

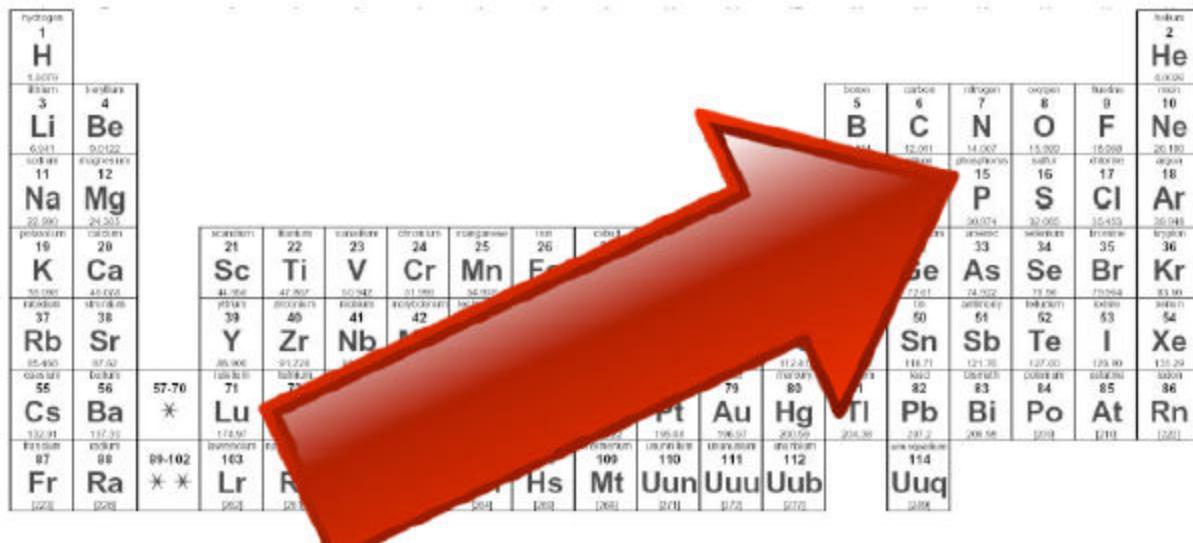
i WIZE TIP

General order of EN may come in handy:

F > O > N > Cl > Br > I > S > C ~ H

The Periodic Trend

Electronegativity increases going up and to the right in the periodic table.



Which atom is the most electronegative? _____

3.24.2

Example: Increasing Electronegativity

Rank the following atoms in increasing electronegativity: O, Al, Sr, N, Si, Cs

Group →	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
↓ Period																		
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																		
Actinides																		
	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			
	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			

Practice: Ranking Electronegativity

Choose the option which correctly ranks the following elements in terms of electronegativity.

Group →	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
↓ Period																		
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			

Ti < Bi < Hg < Cs < Fr

Na < Al < O < S < Cl

Mg < P < C < O < F

Fe < Co < Mo < F < Cl

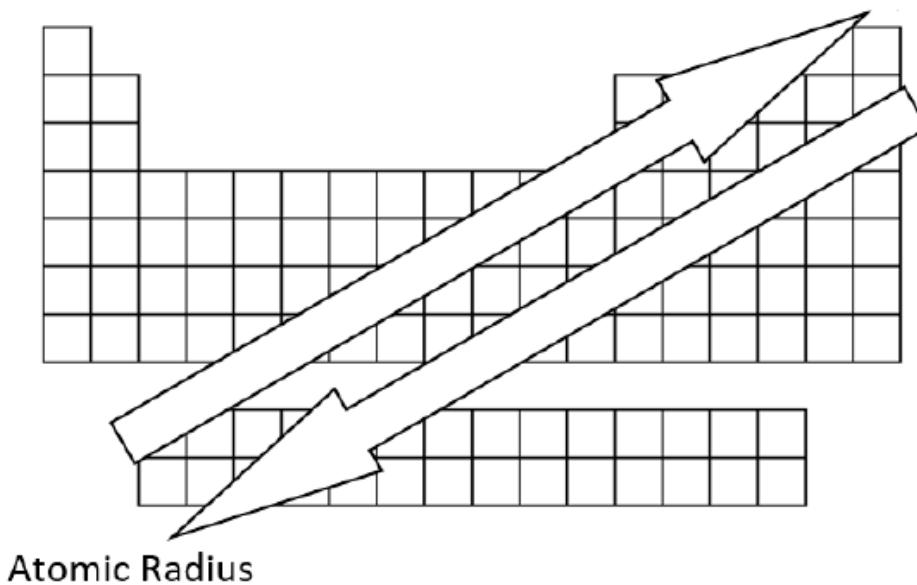
3.25

Summary of Periodic Table Trends

3.25.1

Summary of Periodic Trends

Electronegativity, Ionization
Energy, Electron Affinity



Review: which direction does Zeff increase in? _____

3.25.2

Example: Periodic Trends

Label the following statements as either TRUE or FALSE

1. Ionization energy decreases when the atomic size decreases
2. As atomic size increases it gets easier to add an additional electron
3. Mg^{2+} is the same size as Ne since they are isoelectronic

Practice: Periodic Trends

Complete the statements below by filling in the blanks and then choose which option below would help explain your answer.

Group →	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
↓ Period																		
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			

Part 1

N^{3-} has a (larger/smaller) ionic radius in comparison to Si^{4+} because ...

... of the difference in Z^*

... of the difference in n

... it has a higher proton to electron ratio

... it has a higher electron to proton ratio

Practice: Periodic Trends

Complete the statements below by filling in the blanks and then choose which option below would help explain your answer.

Group →	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
↓ Period																		
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			

Part 2

Al has a (more negative/ less negative) EA than Si because ...

... of the difference in Zeff

... of the difference in n

... it has a higher proton to electron ratio

... it has a higher electron to proton ratio

3.25.4

Practice: Periodic Trends

Which of the following statements is correct about chlorine and phosphorus atoms?

Group →	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18			
↓ Period																					
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) The electron affinity of a chlorine atom is smaller than that of a phosphorus atom due to its smaller radius.
- B) A chloride anion has higher electron affinity than a neutral chlorine atom.
- C) A chlorine atom can easily gain and/or lose an electron since it has both a large electron affinity and a large ionization energy.
- D) The effective nuclear charge experienced by the valence electrons in a chlorine atom is smaller than in a phosphorus atom.
- E) All of the above statements are false.

3.25.5

Practice: Attracting Electrons

For an S-O bond, which atom will attract the electrons within the bond more strongly? Why?

Group →	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
↓ Period																		
1	H																He	
2	Li	Be															F	Ne
3	Na	Mg															Cl	Ar
4	K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
5	Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
6	Cs	Ba		Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
7	Fr	Ra		Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Cn	Nh	F1	Ms	Lv	Ts	Og
Lanthanides	La	Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu			
Actinides	Ac	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr			

Sulfur, because it has a higher electronegativity

Oxygen, because it has a higher electronegativity

Sulfur, because it has a higher Zeff

Oxygen, because it has more valence electrons

Sulfur, because it has more protons

Oxygen, because it is smaller

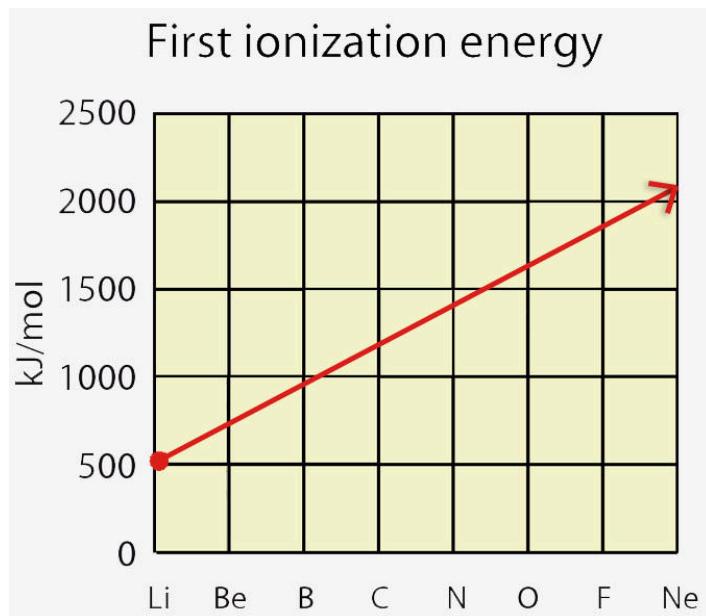
3.26 Exceptions to the Ionization Energy Trend

3.26.1

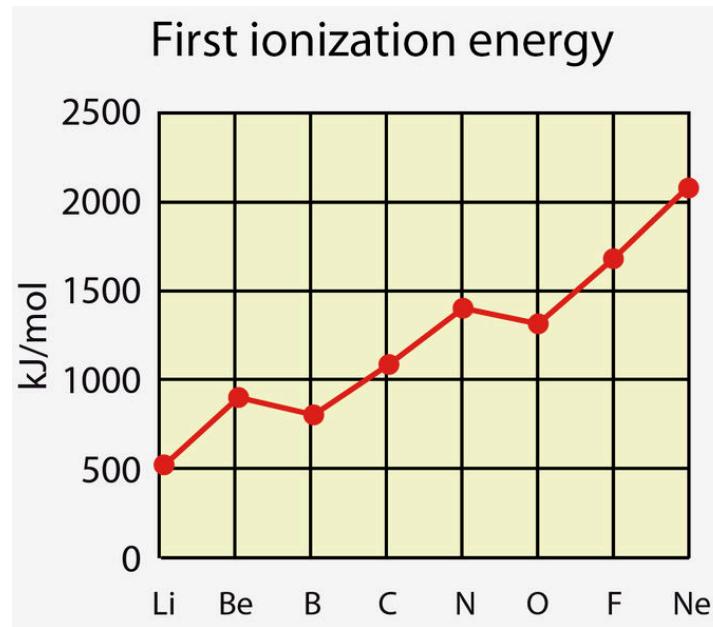
Exceptions to the Periodic Trends Introduction

There are a few exceptions for ionization energy and electron affinity that we should familiarize ourselves with.

Let's consider ionization energy first. **Based on what we currently know**, the trend for ionization energy should look like this:



However, it actually looks something like this:

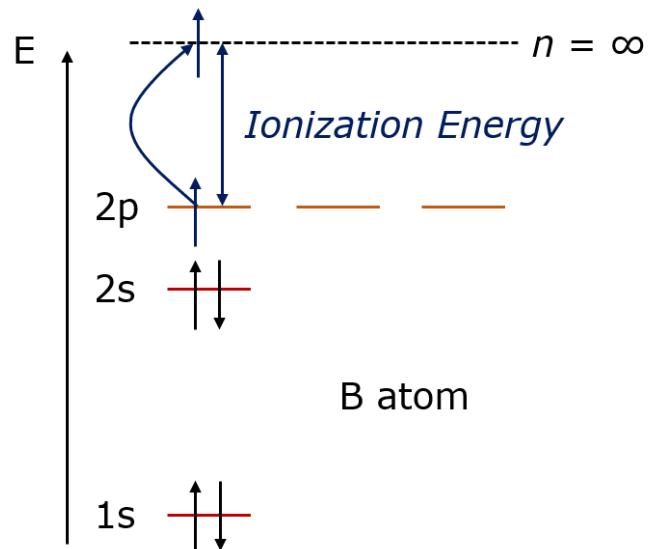
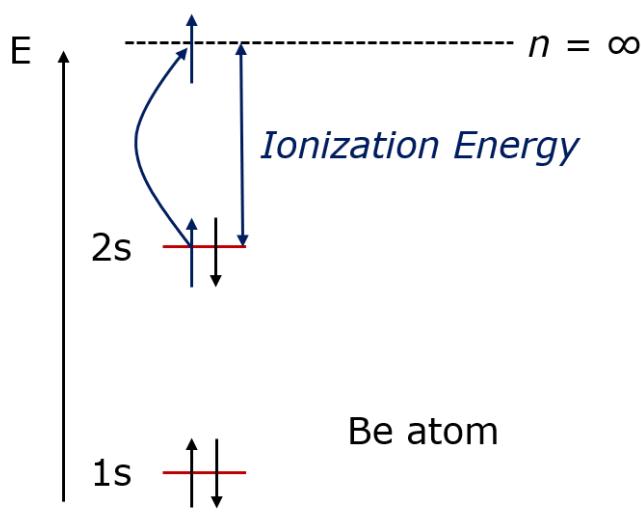


The differences have to do with **electron configurations**. Let's take a look.

Exceptions to the Periodic Trends Examples

Example #1:

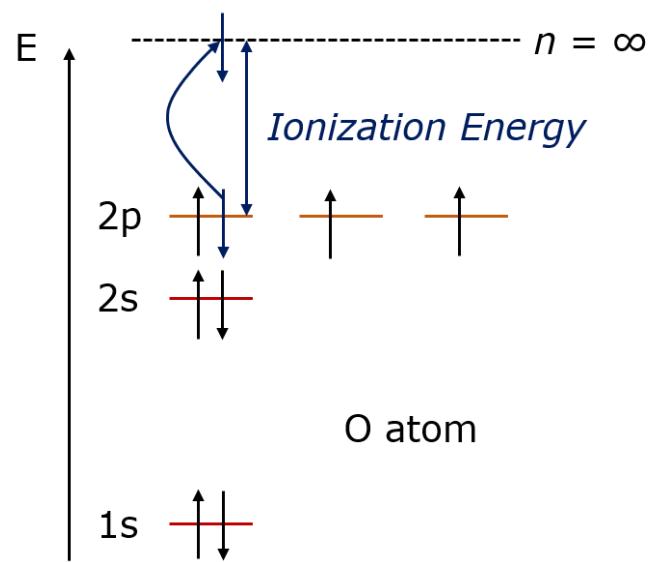
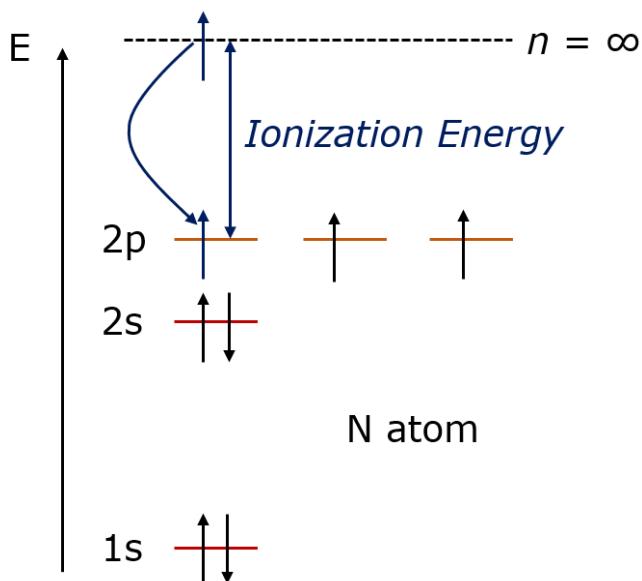
- Based on the periodic trend we learned so far, we would expect that Be or B would have the higher IE: _____



- Recall: which orbital is higher in energy, s or p? _____
- Based on the electron configurations and knowing that **B would LOVE to have its high energy p electron removed** so it can have its **valence electrons in the more stable 2s orbital**, should Be or B have the higher IE? _____

Example #2:

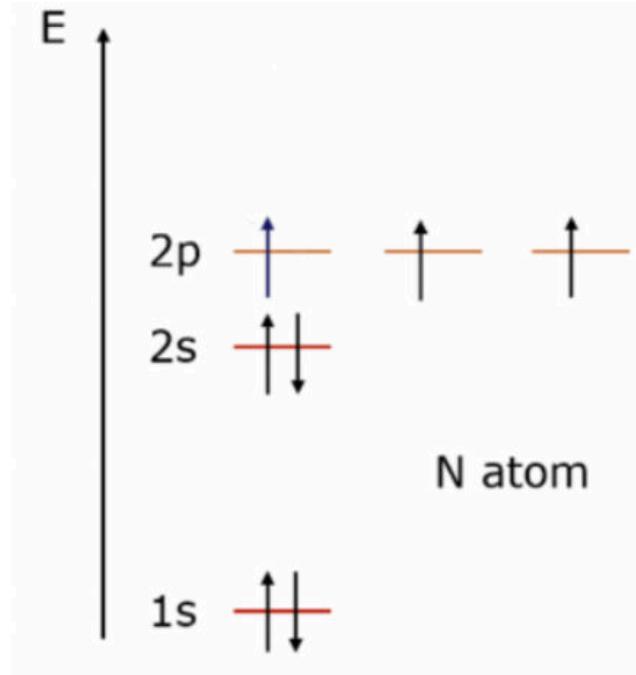
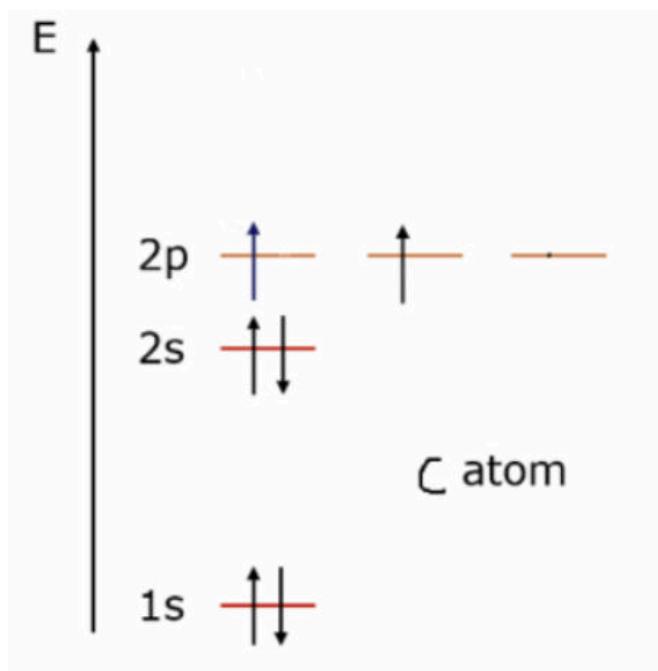
- Based on the periodic trend we learned so far, we would expect N or O to have the higher IE:



- Recall: are half-filled subshells more/less stable than partially filled subshells? _____
- Based on the electron configurations and knowing that **N would HATE to have an outermost electron removed** since **it is currently very stable**, should N or O have the higher ionization energy? _____
- In addition, O has some electron-electron repulsion that it would prefer to get rid of. Thus, it makes sense that it has a higher/lower _____ IE

Example #3:

- Based on the periodic trend so far, we would expect C or N to have the higher negative EA:



- Based on the electron configurations and knowing that **N would HATE to have an electron added since it is currently very stable with a half-filled 2p subshell**, should N or C have the higher electron affinity? _____

Exceptions to the Periodic Trends Summary

WIZE TIP

The exceptions specifically happen in **two cases**:

- When we **switch from one type of orbital to another**
Example: 2s to 2p; (Be to B is a jump from 2s to 2p)
- When we **start pairing electrons in a sub-shell**
Example: 2p³ to 2p⁴ (N vs O)

These exceptions can be explained (and remembered) using the **relative energies of atomic orbitals (knowing that the p subshell is > in energy than the s subshell)** and the fact that electrons repel one another and this repulsion is highest for electrons sharing an orbital (**electrons would rather be half-filled in a subshell** instead of having some electrons paired and others unpaired)

Example: Ionization Energy Exceptions

According to ionization energy (IE) trends, you would expect sulfur to have a higher IE than phosphorus; however, the actual ionization data shows the opposite: the IE of sulfur is 1000 kJ/mol and the IE of phosphorus is 1012 kJ/mol. Explain this phenomenon.

3.26.5

Practice: Ionization Energy Exception

Contrary to what we would predict based on Z_{eff} , the ionization energy of aluminum is actually less than magnesium. This is because,

The 3p orbitals of Magnesium are higher in energy than the 3p orbitals of Aluminum

The 3s orbitals of Aluminum are higher in energy than the 3s orbitals or Magnesium

The 3p orbital of Aluminum is higher in energy than the 3s orbital of Magnesium

The 3s orbital of Magnesium is higher in energy than the 3p orbital of Aluminum

The 3p orbital of Magneisum is higher in energy than the 3s orbital of Aluminum

3.27 The Periodic Table & Properties

3.27.1

The Periodic Table of Elements

The **periodic table** is something we are going to see a lot of in chemistry! It organizes the elements by their **atomic number (Z)** and is organized into **groups (columns)** and **periods (rows)**.

Elements in the **same group have very similar reactivity** which we will talk about more when we learn about things like valence electrons, bonding, and Lewis structures.

Group →	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
↓ Period																		
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			

You should be familiar with each of the following labels:

- Groups
- Periods
- Alkali Metals
- Alkaline Earth Metals
- Transition Metals
- Nitrogen group (aka Pnictogens)
- Oxygen group (aka Chalcogens)
- Halogens
- Noble (Inert) Gases
- Metals
- Non-metals
- Metalloids
- Lanthanides and Actinides (aka Rare Earth Metals)

3.27.2

Practice: Understanding Noble Gases

Which of the following statements is true?

Noble gases are highly reactive since they want to obtain a full octet.

Noble gases are highly reactive since they want to gain more electrons.

Noble gases are unreactive since they already have a full octet.

Noble gases are very stable because they want to gain more electrons