

Unit 3 - Electronic Structure SG

Electrons and Light

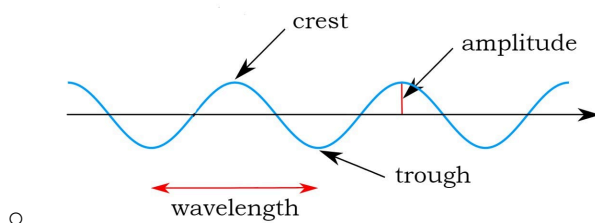
Properties of Light

- All waves travel at the constant speed of light, and light transports energy in the form of a wave
- Frequency (ν) = # of waves that pass through a point in one second
 - Symbol = ν
 - Unit is the Hertz (cycles/second), or s^{-1}
- Wavelength and frequency are **inversely proportional**
 - Low wavelength \rightarrow high frequency
 - high frequency \rightarrow low wavelength

$$c = \nu \lambda$$

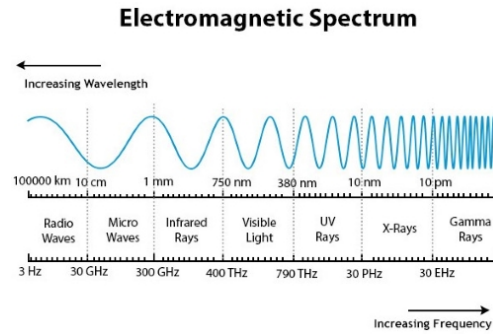
↑ "nu" ↑ "lambda"

- Energy and frequency are **directly proportional**
 - $E = h\nu$
 - $h = 6.63 \times 10^{-34}$ Joules seconds (Planck's constant)
- Energy and wavelength are **inversely proportional**
 - $E = h\nu = hc/\lambda$
 - Increased wavelength \rightarrow decreased energy
 - Decreased wavelength \rightarrow increased energy
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The Electromagnetic Spectrum

- As you go to the RIGHT on the spectrum...
 - Wavelength DECREASES
 - Frequency INCREASES
 - Energy INCREASES
- Order of the EM Spectrum



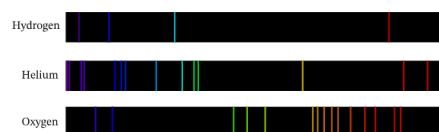
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- **INCREASING FREQUENCY!!!!**
 - Infrared, then visible light, then UV. Infrared has a **LONGER** wavelength.
- Visible light can produce wavelengths of the rainbow - ROY G BIV
 - Red (longest), orange, yellow, green, blue, and violet (shortest)

Coulombic Attraction

- Coulombic attraction is the attraction between oppositely charged particles
 - Protons in nucleus attract electrons surrounding it due to their opposite charges
- Force of attraction is **INVERSELY PROPORTIONAL** to distance
 - As the distance of the outermost electron from the nucleus increases, the attractive force decreases.
- Periodic Trends
 - **Down a Group/Column:** force of attraction **DECREASES**
 - Because distance of outermost electron to the nucleus **INCREASES**
 - While increasing the # of protons strengthens attraction, the **distance** of the outermost electron has a larger impact on the force
 - **Across a Period/Row:** force of attraction **INCREASES**
 - Because # of protons increases
- The attractive force to the nucleus (nuclear attractive force) is experienced equally by all electrons in the atom, regardless of the number of electrons.

Electron Energy and Light

- Background
 - Giving electrons energy leads to light
 - Highest wavelength & lowest energy = Red
 - Lowest wavelength & highest energy = Violet
- The spectral lines for each atom are their fingerprints



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- Electron Gaining and Losing Energy
 - Moving closer to nucleus - **LOSE** energy
 - Moving further away from nucleus - **GAIN** energy
 - As an electron gets closer to the nucleus, the attraction gets stronger. For an electron to move further from the nucleus it would need to **GAIN** energy because it moves **AGAINST** the attractive force.
 - Another way to look at it is electric PE: $U = (-k) \cdot q_1 \cdot q_2 / r^2$, hence as r increases, U becomes less negative and also **increases**.
- Ground and Excited States
 - Light interacted with electrons to produce spectral lines.
 - Electrons **absorb** energy when moving from **ground** states (lower energy levels, lower PE) to **excited** states (higher energy levels, higher PE)
 - If any ONE ELECTRON is not at the ground state, then the entire atom is excited
 - Example: ground state of Hydrogen - 1s. **Excited**: 2s, or 2p, etc.
 - Low->high: absorb energy; high->low: release energy
- When moving from a high to low energy state, electron **releases energy in the form of LIGHT**
 - More energy levels fallen -> more energy is released -> smaller wavelength
 - The wavelength corresponds to the color on the color spectrum
 - If you fall only **ONE** energy level, you will have the **longest** wavelength, hence lowest energy
 - If you fall the **MOST** energy levels, you will have the **shortest** wavelength, hence highest energy
 - A single atom of hydrogen cannot produce all four spectral lines at once since you only have one electron. You can make it appear to emit all colors at the same time by having lots of hydrogen atoms, which could have different energy transitions (e.g. one of $n=3$ to $n=2$, one $n=4$ to $n=2$, etc.)
- Bright line spectra (the **UNKNOWN**)
 - Emissions spectra from multiple elements have lines of each element (mixed colors)
 - If there is ONE LINE that does not have a correspondence in the UNKNOWN, that completely disqualifies them from being in the Unknown
- Ground states
 - Lower energy levels (closer) -> higher (further) -> absorb energy
 - Electrons SPONTANEOUSLY drop down to a ground state, emitting energy in the form of **LIGHT**; energy emitted is equal to energy drop of the given distance

Flame Test Lab

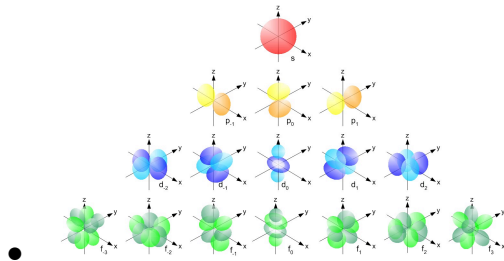
- Colors emitted during a flame test vary because different elements have unique electron configurations. When an element is heated, its electrons absorb energy and jump to higher energy levels. As these excited electrons return to their original (ground) state, they release the absorbed energy in the form of light. The specific wavelength (or color) of this light depends on the difference in energy between the electron's excited state and its ground state.

- Each element has a unique set of energy levels, so the energy gaps (and thus the emitted wavelengths of light) differ, producing different colors in the flame test. For example, sodium emits a yellow flame, while copper gives a green flame due to these unique electron transitions.

Electronic Structure

Quantum Mechanical Models

- Principal Energy Level (n) / also known as shell
 - Describes probable distance of electrons from nucleus
 - Principle quantum number ($n=1,2,\dots,7$)
 - $n=1$ = LOWEST energy level, closest to nucleus
 - $2n^2$ describes the max # of electrons in the nth shell
- Sublevel / subshells
 - Group of 1 or more orbitals of the same energy level (e.g. 1) and usually the same shape
- Orbital
 - Region within an energy level where there is a probability of finding an electron
 - Orbital shapes = defined as the surface that contains 90% of total electron probability
 - Each orbital / shape holds a max of 2 electrons
- Types of Orbitals
 - S orbital (sharp) - Spherical, centered at origin
 - First appears at the first energy level (1s); max 1 orbital per energy level
 - Hence, max of 2 electrons per orbital
 - P orbital (principal) - Peanut shape
 - First appears at the 2nd energy level (2p), max 3 orbitals per energy level
 - D orbital (diffuse) - double dumbbell/dumbbell with donut
 - First appears at 3rd energy level (3d), max 5 orbitals per energy level
 - F orbital (fundamental) - funky/flowery
 - First appears at 4th energy level, max 7 orbitals per energy level



Energy Level (n)	Sublevels	# orbitals/sublevels (orientations)	# of electrons/sublevels	Total electrons for each energy level
1	s	1	2	2
2	s, p	1+3=4	2+6	8

3	S, p, d	$1+3+5=9$	$2+6+10$	18
4	S, p, d, f	$1+3+5+7=16$	$2+6+10+14$	32

- Max orbitals per sublevel

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Sublevel	Max # of orbitals	Max # of electrons
s	1	2
p	3	6
d	5	10
f	7	14

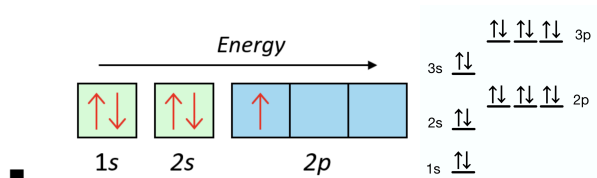
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- Electron Spin

- Each orbital (orientation) is only allowed TWO electrons
 - Electron spin describes behavior/direction of spin of an electron in a magnetic field

Orbital Diagrams and Electron Configuration

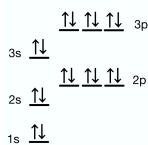
- You should know **who's principle said what**, and also whose rule was broken given an orbital diagram
- Orbital diagram
 - Visual representation of orbitals in an atom and shows order that orbitals are filled in (from lowest->highest energy)
 - Every "box" represents an orbital



- Orbitals @ the same sublevel have the same energy
 - No more than 2 electrons / arrows on a line

- Aufbau Principle**

- An electron occupies the **lowest energy orbital** it can...



- fill arrows from the bottom up

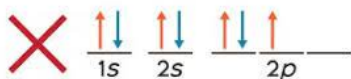
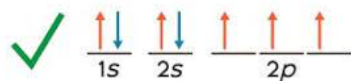
- Pauli Exclusion Principle**

- No two electrons in the same atom can have the same position and spin

- If there are 2 electrons in an orbital, 1 spins up and 1 spins down (1 up and 1 down arrow)
- No 2 electrons can have the same **ADDRESS**

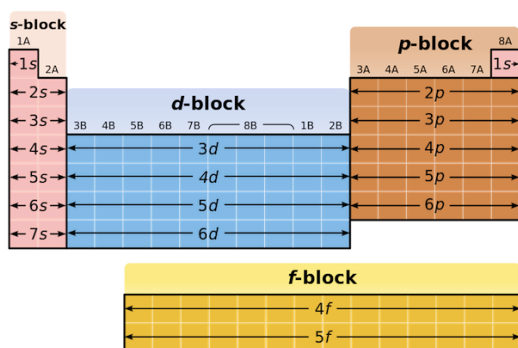
- **Hund's Rule**

- Orbitals of **equal energy** (same sublevel) are occupied by one electron before any orbital is occupied by a second electron
 - I.e. electrons do not want to share unless they have to
- If more than one orbital at the same energy sublevel (s,p,d,f etc.), then put one electron into each orbital BEFORE pairing up.
- First arrows can be UP or DOWN. But they must be the same spin in the same sublevel
- You can also fill 3rd box before 2nd before 1st etc. but it's unconventional



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- **Electron Configuration**
 - Shorthand notation for expressing filled orbitals (location of electrons) of an atom
 - Three components - energy level, sublevel (s, p, d, f), # of electrons
 - E.g.: H would be $1s^1$ at ground state, where the 1 in exponent symbolizes number of electrons
 - Carbon would be $1s^2 2s^2 2p^2$
 - Each energy level (1,2,...,7) has sublevels (s,p,d,f)
- **Paired and unpaired electrons**
 - **Paired** = 2 electrons in a box, **unpaired** = 1 electron in a box
- The test will ask you: how many paired/unpaired electrons are in this element, etc.
 - Also be prepared to identify the element given the electron configuration

Periodic Table Shortcut & Noble Gas Notation Notes



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- The periods/rows are the principal energy levels (1,2,3,...,7)
- Down a column, you have similar electron configurations

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- Notice how in the p-block, you have 6 columns/families which is equal to the maximum # of electrons that the p-sublevel can hold in its 3 orbitals.
- Based on where an element is on the periodic table you can easily determine which orbital and energy level the last electron is filled
 - **Example:** Oxygen (O): Period 2, Group 16 → Last electron fills a 2p orbital.
- Notice:
 - S and p are always matching with row and period #
 - D block rows are always 1 energy level behind the period they are listed in ($n = \text{period} - 1$, so order $4s \rightarrow 3d \rightarrow 4p$)
 - F block rows are always 2 energy levels behind the period they are listed in ($n = \text{period} - 2$, so order $6s \rightarrow 4f \rightarrow 5d \rightarrow 6p$)
- Order of filling
 - 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s, 5f, 6d, 7p
- Noble Gas Notation
 - Find the element on the p-table, and the **noble gas** in the **PREVIOUS** period
 - Write the noble gas symbol inside brackets
 - List any remaining electron configuration left over until you get to your element

Electron Configuration Exceptions & Ion Formation

- a D or a F subshell is particularly stable if it is half or completely full
- $s^2 d^4 \rightarrow s^1 d^5$, and $s^2 d^9 \rightarrow s^1 d^{10}$
- Exceptions to the electron configuration

Unusual Electron Configurations

Element	Predicted Electron Configuration	Actual Electron Configuration
copper, Cu	[Ar] $3d^9 4s^2$	[Ar] $3d^{10} 4s^1$
silver, Ag	[Kr] $4d^9 5s^2$	[Kr] $4d^{10} 5s^1$
gold, Au	[Xe] $4f^{14} 5d^9 6s^2$	[Xe] $4f^{14} 5d^{10} 6s^1$
palladium, Pd	[Kr] $4d^8 5s^2$	[Kr] $4d^{10}$
chromium, Cr	[Ar] $3d^4 4s^2$	[Ar] $3d^5 4s^1$
molybdenum, Mo	[Kr] $4d^4 5s^2$	[Kr] $4d^5 5s^1$

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- Know Cr, Mo, Cu, Ag

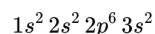
- You do NOT have to know the f-block exceptions
- Ion Formation
 - Atoms gain or lose an electron during chemical reactions in order to try to achieve a noble gas configuration, because noble gases have very stable electron configurations with all filled orbitals, and so are very unreactive

+1															0
H ⁺	+2									+3	+/-4	-3	-2	-1	He ⁰
Li ⁺	Be ²⁺									B ³⁺	C	N ³⁻	O ²⁻	F ⁻	Ne ⁰
Na ⁺	Mg ²⁺									Al ³⁺	Si	P ³⁻	S ²⁻	Cl ⁻	Ar ⁰
K ⁺	Ca ²⁺									Ga ³⁺	Ge	As ³⁻	Se ²⁻	Br ⁻	Kr ⁰
Rb ⁺	Sr ²⁺									In ³⁺	Sn	Sb ³⁻	Te ²⁻	I ⁻	Xe ⁰
Cs ⁺	Ba ²⁺									Tl ³⁺					Rn ⁰
Fr ⁺	Ra ²⁺														

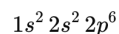
- (memorize)
- Main group elements form ions periodically
- NOTE: for the test, when a problem says **ion**, do the electron configuration for the ion (e.g. if Lithium ion, do e-config for Lithium 1+ by removing 1 electron)
- When it says an **ATOM**, that means do the electron config for the neutral atom
- Atoms gain or lose electrons in their HIGHEST ENERGY LEVEL.
 - Write the e-config of neutral atoms first, then add/subtract electrons
 - Note that the # of protons and neutrons do NOT change

The electron configuration for a **magnesium ion (Mg²⁺)** is derived by removing two electrons from the neutral magnesium atom:

- **Neutral Magnesium (Mg)**: Atomic number 12, with electron configuration:

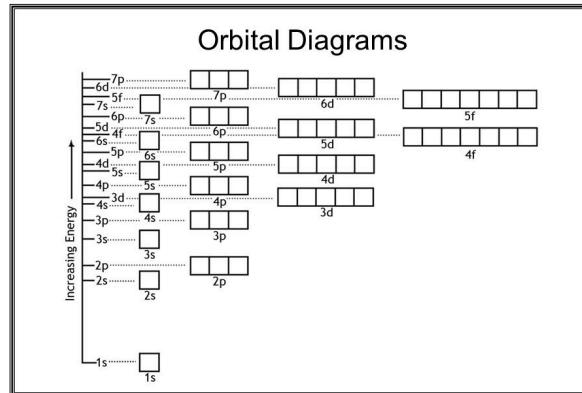


- **Magnesium Ion (Mg²⁺)**: To form Mg²⁺, magnesium loses its two outermost electrons from the 3s orbital:



The electron configuration for **Mg²⁺** is the same as **Neon (Ne)**, making it **1s² 2s² 2p⁶**. This is a stable, noble gas configuration.

- Example:
- Electron LOSING order is **DIFFERENT** from the electron filling order (which always goes in order)
 - Electrons are lost from the highest principal energy level first (outermost shell)
 - Once you remove all of the highest energy level and still need to remove, you remove from highest energy sublevel available within the next level
 - In general if they have the same number before (like 3p and 3s), you remove f, before d, before p, before s



And Electron Configurations

Gallium's Electron Configuration Recap

- Neutral Gallium: $[\text{Ar}] 4s^2 3d^{10} 4p^1$

To form Ga^{4+} , we need to remove **four electrons**. Here's the process:

1. Remove electrons from the outermost energy level first ($n=4$):

- Remove the **4p electron** first, then the **two 4s electrons**.
- After removing these three electrons, the configuration becomes:
 $[\text{Ar}] 3d^{10}$
- At this point, we have Ga^{3+} .

2. Next, remove from the highest available sublevel in the next outermost level ($n=3$):

- The next electron to remove would indeed come from the 3d sublevel rather than 3s or 3p because 3d is **higher in energy than both 3s and 3p**, even though they're in the same principal level.
- This removal gives us Ga^{4+} with the configuration:



Example:

Quantum Numbers

Type	Symbol	Allowed values	Property of orbital
Principal Quantum #	n	1, 2, 3, ...	Determines shell size and energy (shell number)
Angular Momentum quantum #	ℓ	0 to $n-1$	Shape of orbital ($\ell=0 \rightarrow$ s-sublevel, $\ell=1 \rightarrow$ p-sublevel, $\ell=2 \rightarrow$ d-sublevel, $\ell=3 \rightarrow$ f-sublevel)
Magnetic quantum #	m_ℓ	$-\ell$ to $+\ell$	Spatial orientation of orbitals / # of orbitals @ your sublevel
Electron spin quantum #	m_s	$-\frac{1}{2}$ and $+\frac{1}{2}$ (for convention, use up + arrows before down -)	Spin state of electron

- You write quantum #s in the form (n, ℓ, m_ℓ, m_s)
- Principal quantum #
 - Example: 1s, 4d, 7p

- Angular Quantum #

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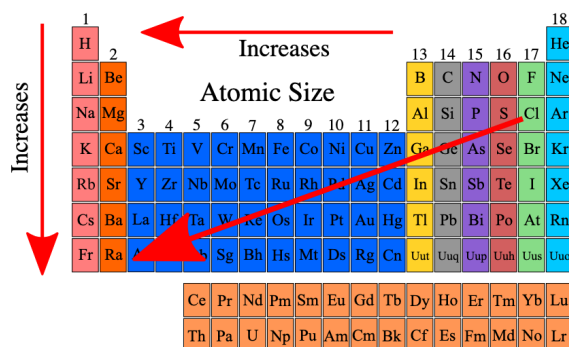
ℓ	sublevel
0	s
1	p
2	d
3	f

- If $n=2$, then ℓ can take on values from 0 to 1, so $n=2$ can have the s and p orbitals
 - Example: first two quantum #s (n, ℓ) of 3d would be $n=3$, and $d \rightarrow \ell=2$, so (3,2)
 - ℓ depends on the sublevels
- Magnetic Quantum #
 - Example: first 3 quantum numbers for 3d electron of Scandium: $n=3$, $d \rightarrow \ell=2 \rightarrow m_\ell = \{-2, -1, 0, 1, 2\}$. You could have (3,2,-2), (3,2,-1), (3,2,0), (3,2,1), (3,2,2)
- Electron spin quantum #
 - Example: the 4 quantum numbers for Germanium's 31st electron: $\text{Ge} = 4p^2$, so 31st = $4p^1$. $n=4$, p sublevel, hence $\ell=1$ hence $m_\ell = \{-1, 0, 1\}$. Conventionally you spin up ($+\frac{1}{2}$)
 - If 31st electron is (4,1,0, $+\frac{1}{2}$), then 34th electron has to spin DOWN but can be in any of the 3 positions since it's the first spin down after all the up arrows are filled.
- Magnetic Quantum Numbers, continued
 - $\ell=0$ (s-sublevel): $m_\ell = \{0\}$, only one total orbital/box
 - $\ell=1$ (p-sublevel): $m_\ell = \{-1, 0, 1\} \rightarrow 3$ total orbitals/boxes/lines
 - $\ell=2$ (d-sublevel): $m_\ell = \{-2, -1, 0, 1, 2\} \rightarrow 5$ total
 - $\ell=3$ (f-sublevel): $m_\ell = \{-3, -2, -1, 0, 1, 2, 3\} \rightarrow 7$ total
 - The numbers in brackets represent the possible addresses for orbitals
- Pauli exclusion principle: one is spin up (+), and one is spin down (-). Conventionally, use (+) before (-).
- How to figure out if something is spin up or spin down
 - If you have the p-sublevel and you want to find the possible quantum numbers of last electron of Sulfur (16 on p-table), it ends in $3p^4$ so 4 electrons. However p-sublevel only has 3 orbitals, so all the previous electrons must be spin-up (conventionally, you could also have spin-down but conventionally do spin-up. Do note that it is NOT ambiguous and you can only have one option for spin if you are given the previous quantum #s) Hence your last electron must be spin down due to hund's rule. You can put this anywhere as you have not placed any spin downs yet.
 - If you want to find location of second to last electron of sulfur ($3p^3$) and are given that the third to last and fourth to last have quantum numbers (3,1,0, $+\frac{1}{2}$) and (3,1,1, $+\frac{1}{2}$). Then your second to last electron MUST have spin $+\frac{1}{2}$ due to pauli exclusion and must be in -1 due to hund's rule

Periodic Table Trends

- Atomic Radius

- Measure of the size of an atom. Measured as half the distance between nuclei in a bonded diatomic (2-atom) molecule
- Across a period: radius **DECREASES**
 - More protons → greater coulombic attraction → electrons pulled closer to nucleus
- Down a group: radius **INCREASES**
 - Atoms get bigger from adding energy levels/shells (# of energy levels corresponds with the period #)
 - Distance is the bigger deal, so although there is more protons too, the increase in radius is more significant
- Francium would be the largest atom. Helium would be the smallest atom

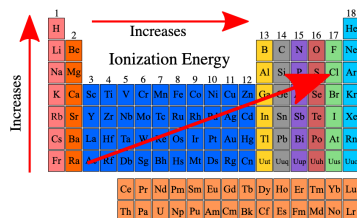


- Ionic Radii

- Cations (+): smaller than the neutral atom (because you lose electrons, the electrons can gather closer to the nucleus, each experience bigger pull)
- Anions (-): larger than the neutral atom, gain electrons

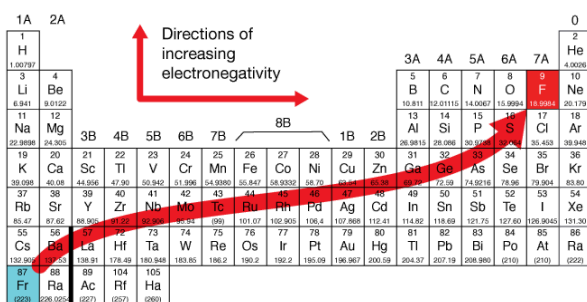
- First** Ionization Energy (IE)

- When referred to as just Ionization energy, that would be the **FIRST** ionization energy i.e. the amount of energy required to remove the 1st outermost electron
- Amount of energy required to remove an outer electron of an atom to form a cation
- Across a period: IE **INCREASES**
 - Because more energy is needed to remove an electron as more protons are added to the nucleus
- Down a group: IE **DECREASES**
 - Because the further from electrons that the nucleus is, the weaker the coulombic attraction. Hence, less work is needed to pull an electron out
 - There are also more protons but **distance is the bigger deal**.



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- 2nd and 3rd Ionization Energy
 - Each additional electron removed requires more and more energy (as electrons gather close and closer, the greater the coulombic force of attraction, hence you need to overcome more force)
 - There is significantly **more** energy required to remove a core electron than an outer electron
- Electronegativity (EN)
 - Ability of an atom in a molecule to attract shared electrons to itself
 - Noble gasses tend to have very low / close to 0 electronegativity
 - Across a period: Electronegativity **INCREASES**
 - This is because coulombic attraction goes up, so it is easier to pull other electrons to the nucleus
 - electronegativity increases **UNTIL** you get to the noble gasses where EN is the **LOWEST**
 - Down a group: Electronegativity **DECREASES**
 - This is because there is a weaker coulombic attraction. Hence, nucleus is less able to hold the electrons that are shared
 - Fluorine is the MOST ELECTRONEGATIVE

Electronegativity Trend



Electron Energy and Excited States

- Electrons close to the nucleus are at the ground state
 - If ALL the electrons are at ground state, then the atom is at a ground state.
- When a photon hits an electron to a high energy level, then the whole atom is at an excited state
 - If one electron is excited, then the whole atom is excited
 - Excited - ABSORBS energy
- Electron configuration examples

- Ground state of Helium: $1s^2$
- Excited state example: $1s^1 2s^1$ or $2s^2$, or $1s^1 2p^1$, etc. (infinite possibilities)
 - Excited states will break Aufbau's principle that electrons fill subshells of lowest available energy
 - $2s^1 2s^2$ means two photons hit the electrons to one higher energy level, $1s^1 2p^1$ means one photon hits an electron to one higher energy level but this time to the p sublevel rather than to the s. Infinite possibilities
- If Aufbau's principle is **violated**, then an atom is at an excited state.
 - Sublevels of lesser energy should be filled to not violate his principle
 - If you skip a WHOLE ENERGY LEVEL \rightarrow also violates Aufbau's principle
 - OR if you have an incomplete previous sublevel \rightarrow also violates
- To find the element:
 - Take one electron from the valence electrons to the sublevel with the missing electrons
 - E.g. if you had Helium $1s^1 2s^1$, take the 1 electron from 2s and put it to the sublevel with missing electron (the 1s sublevel). Hence you identify the element as helium ($1s^2$)
 - OR you can count the # of electrons
- Example: Li [He] $3s^1$ violates Aufbau's principle and is thus excited because it skips a whole energy level (2)
- REMEMBER: Cr, Mo, Cu, and Ag are EXCEPTIONS and their ground state configurations do NOT violate Aufbau's principle
- Each element has a set amount of electrons. This will help you differentiate between which element is at the excited state
 - Example: Mo = [Kr] $5s^1 4d^5$
 - Nb = [Kr] $5s^2 4d^3 \rightarrow$ excited [Kr] $5s^1 4d^4$. D only has 4 not 5 so you can tell the difference between whether it's Nb excited or Mo ground
- You should be able to identify diagrams at excited states / violates aufbau's rule
 - Periodic Table families

H																	He
Li	Be									B	C	N	O	F	Ne		
Na	Mg									Al	Si	P	S	Cl	Ar		
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
Fr	Ra	Ac	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Cn	Nh	Fl	Mc	Lv	Ts	Og
Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu				
Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr				

- Alkali metals, alkaline earth metals, transition metals, chalcogens, halogens, noble gases, lanthanides & actinides / inner transition metals

- Metals, Nonmetals, and Metalloids

Metal										Metalloid			Nonmetal				
H																He	
Li	Be											B	C	N	O	F	Ne
Na	Mg											Al	Si	P	S	Cl	Ar
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
Cs	Ba	La-Lu	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
Fr	Ra	Ac-Lr															
			La	Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu
			Ac	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr

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- Most elements exist as metals

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