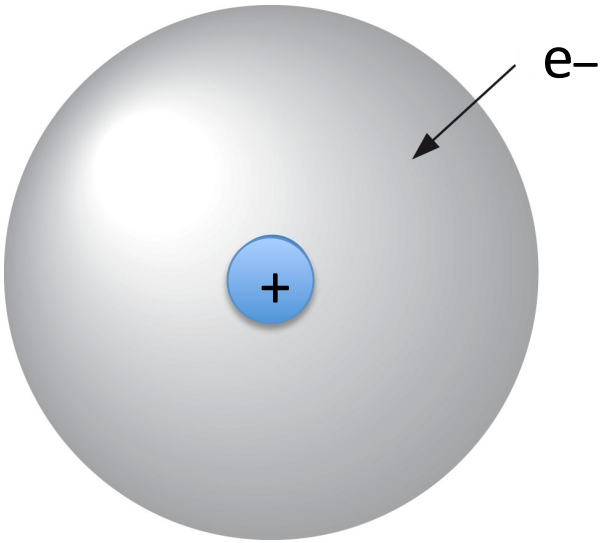
Ch 9. Electrons in Atoms and the Periodic Table

Electronic Structure of Atoms

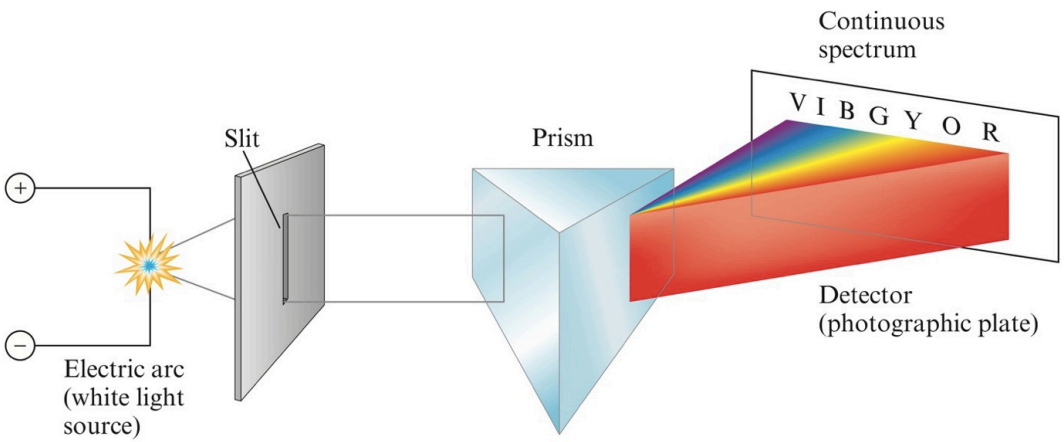
Electronic Structure of Atoms

Rutherford's Nuclear Model of the Atom

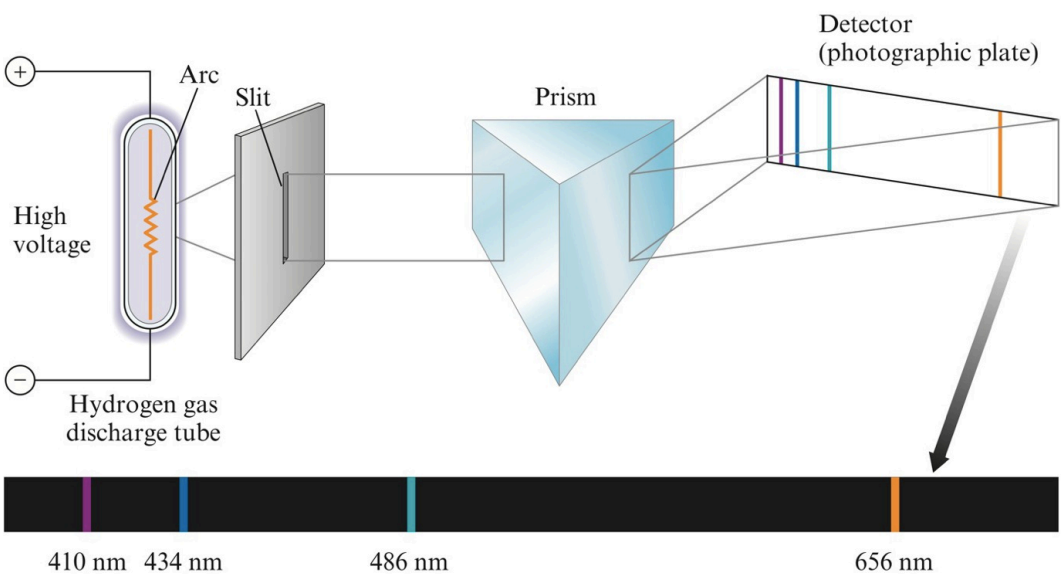


Question: How are the electrons arranged?

Atomic Spectra



White light emits continuous spectrum.



Hydrogen gas emits discrete spectrum (line spectrum).

WHY?

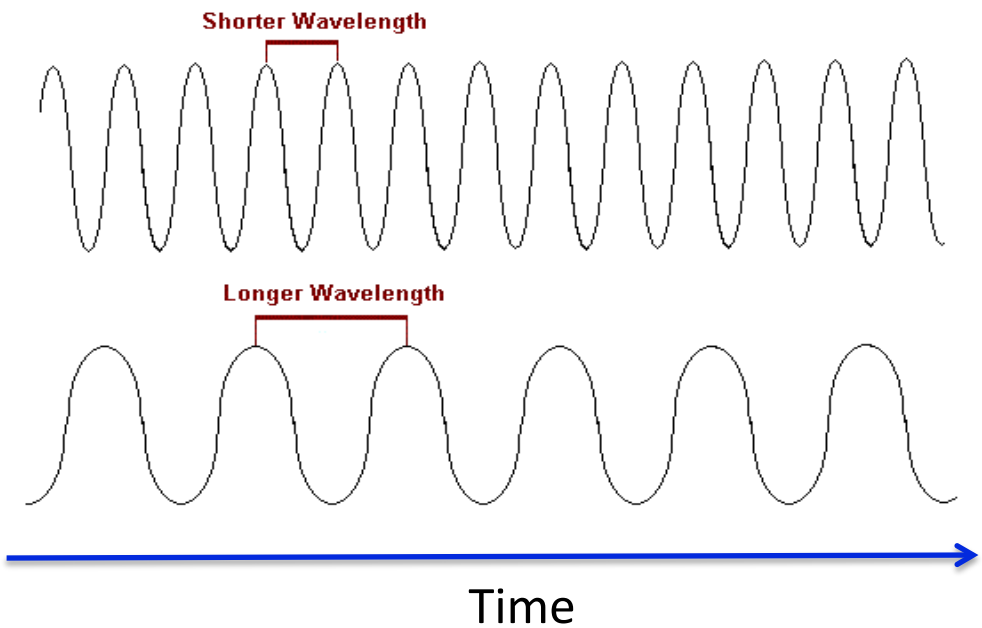
Electromagnetic Radiation

- Electromagnetic Radiation: A form of energy that travels through space through a wave-like behavior.
- Electromagnetic waves are classified by their frequencies (or wavelengths).

Electromagnetic Radiation

- Frequency (ν): number of waves that pass through a given point per time period [#/s].

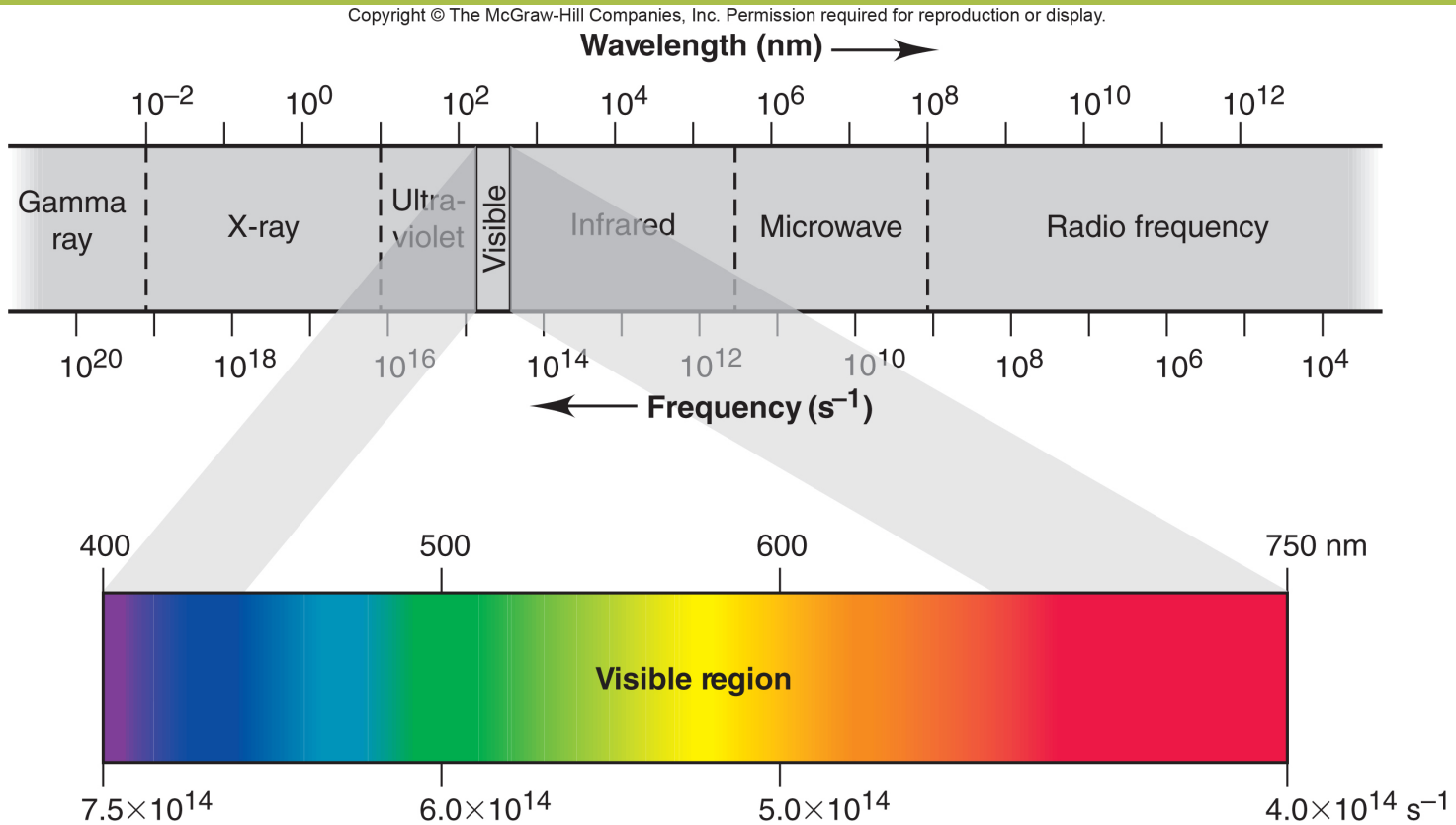
  Higher frequency → Higher energy
- Wavelength (λ): distance between two consecutive wave peaks [nm]
- $c = \lambda\nu$ (c is speed of light, 2.998×10^8 m/s)



Higher Frequency,
Shorter wavelength

Lower Frequency,
Longer wavelength

Electromagnetic Spectrum



Higher Energy
(shorter wavelength, higher frequency)

Lower Energy
(longer wavelength, lower frequency)

Ex Probs

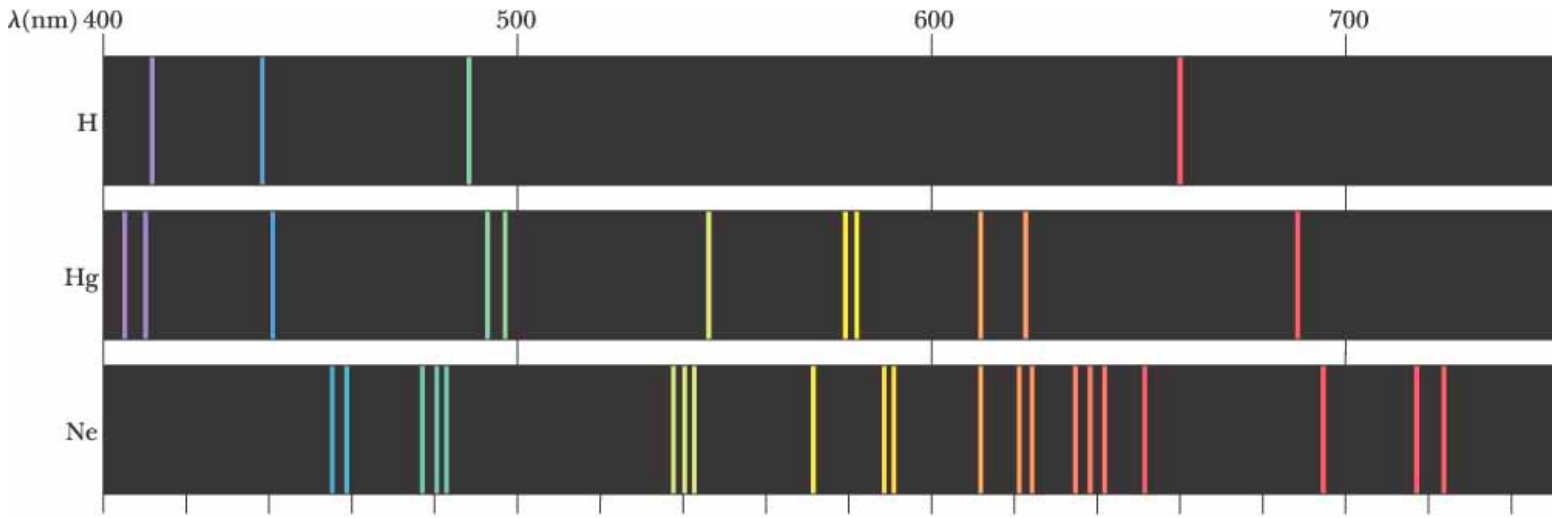
Emission Spectra of Different Atoms

For each type of atom, only certain wavelengths of light are emitted, because…

H

Hg

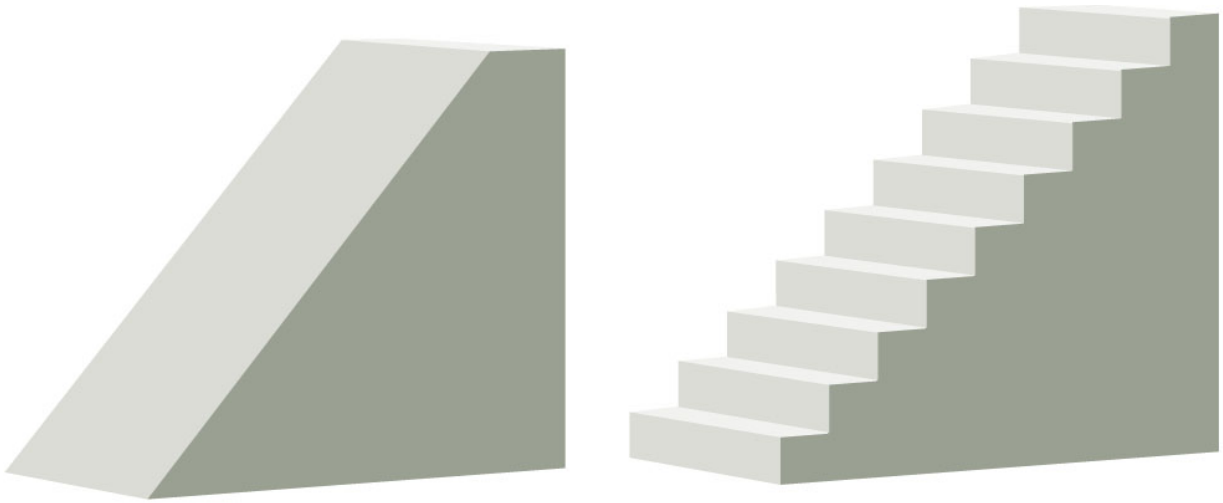
Ne



Energy Levels of Atoms are Quantized

Energy levels of each atom are quantized: Only certain, discrete levels of energy are allowed.

- quantum: a specified portion, a packet

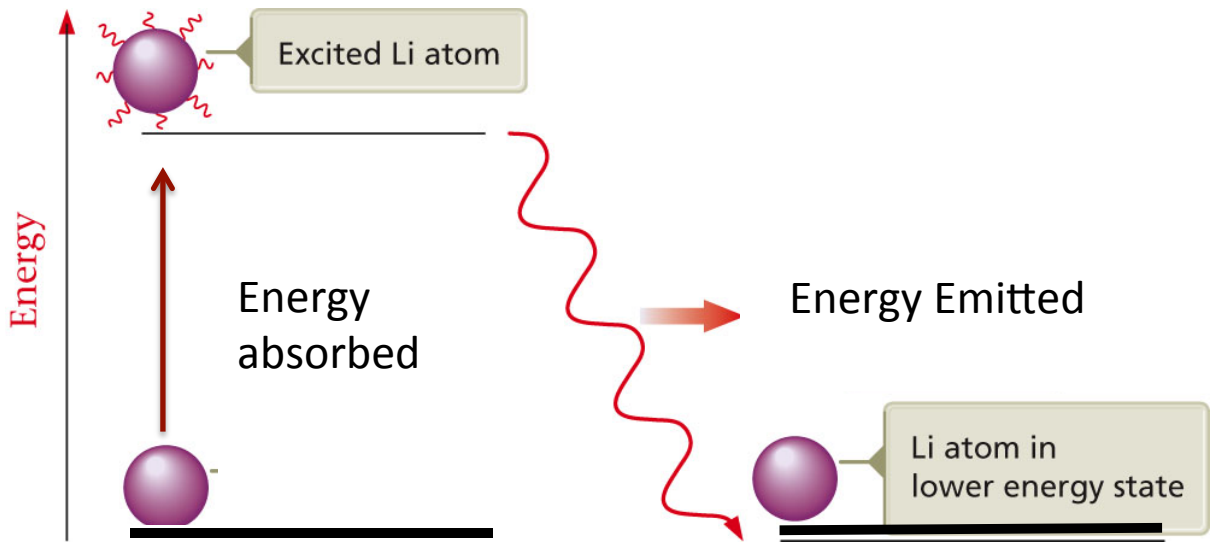


Continuous Quantized

Atoms Absorb/Emit Electromagnetic Radiation (Light)

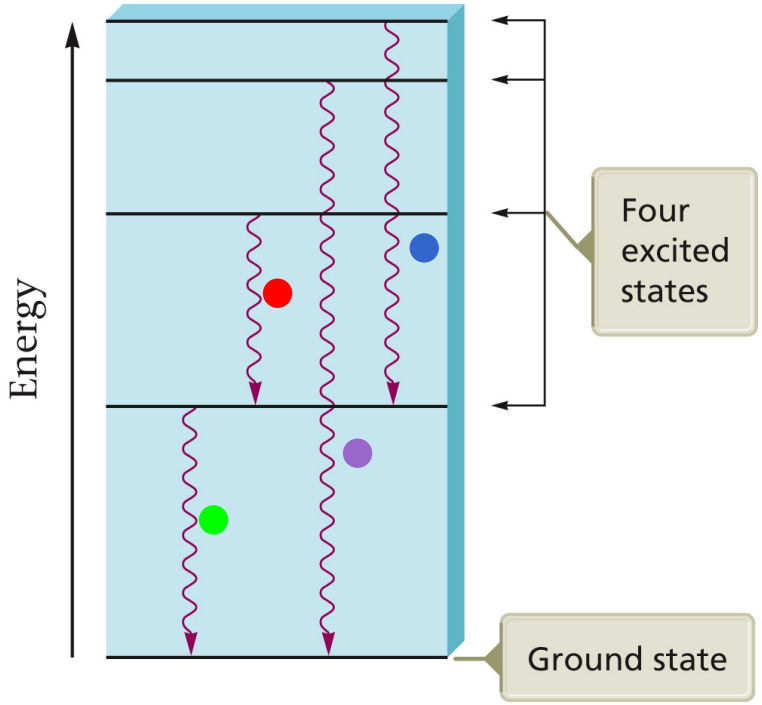
Atoms can exist in different energy levels (states).

1. Atom starts in lowest energy state (ground state).
2. It absorbs energy and becomes excited to a higher-energy state (excited state).
3. As atom goes back down to low-energy state, it releases the energy back: Energy absorbed = energy released



Energy Levels of Atoms are Quantized

Energy Levels for Hydrogen



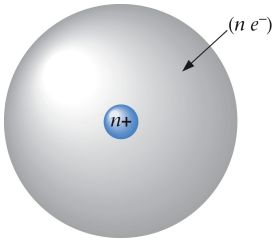
Wavelength emitted depends on size of gap between energy levels (Greater energy gap → greater energy → higher frequency).

Energy Levels of Atoms

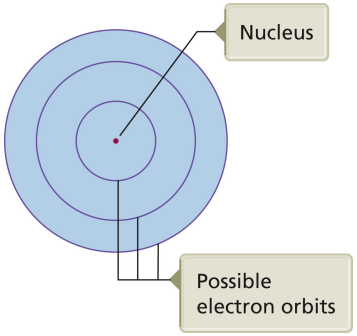
- Energy level of an atom depends on the energy level(s) occupied by its electron(s).

Models of the Atom

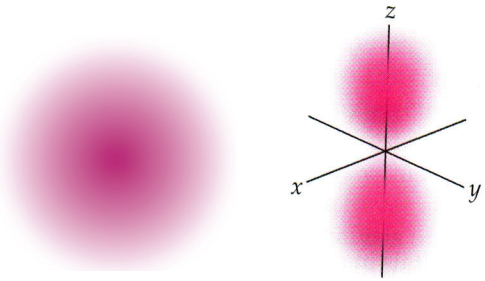
1. Rutherford Model (1911-1932)



2. Bohr Model (1913) incorrect



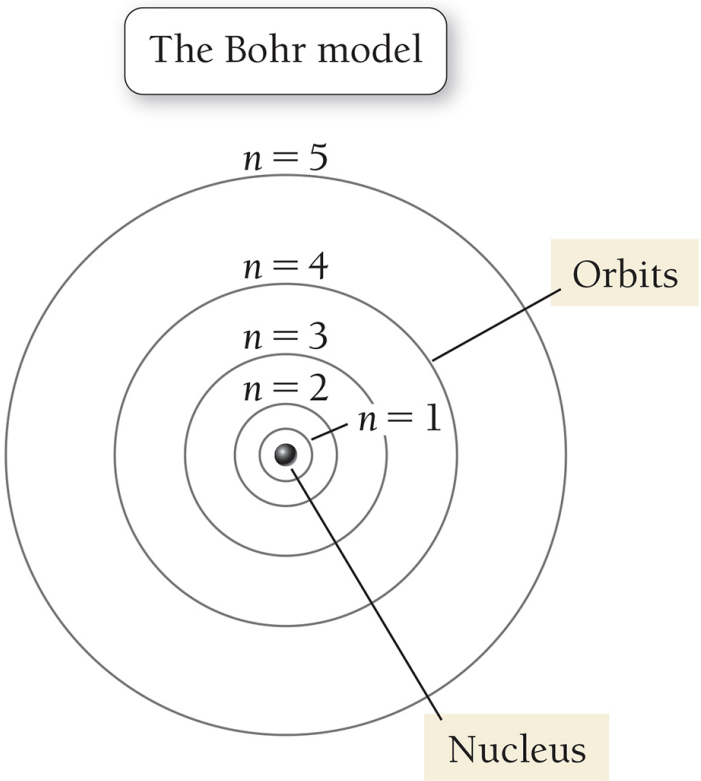
3. Quantum Mechanical Model (1920s): orbitals



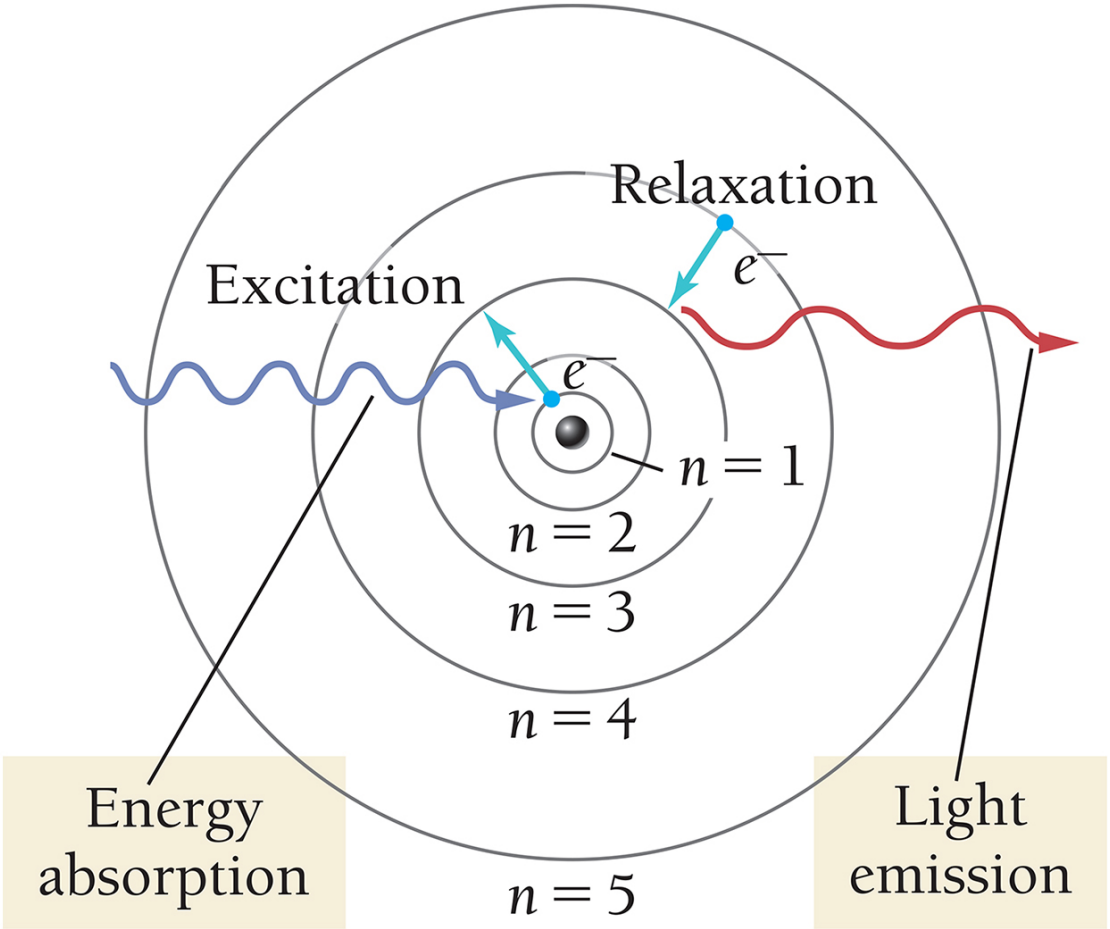
Bohr Model of the Atom (1913)

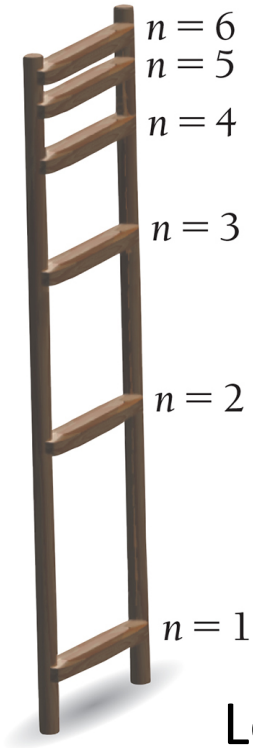
- Electron moves in circular orbits (quantized energy levels).

- Electron jumps between quantized energy levels by absorbing or emitting energy of a particular wavelength → line spectrum



Bohr Model of the Atom





Lowest energy

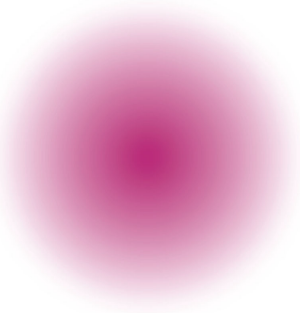
Bohr Model of the Atom

Bohr's model of the atom was incorrect.

- Model does not work for atoms other than hydrogen.
- Electrons do not move in a circular orbit!

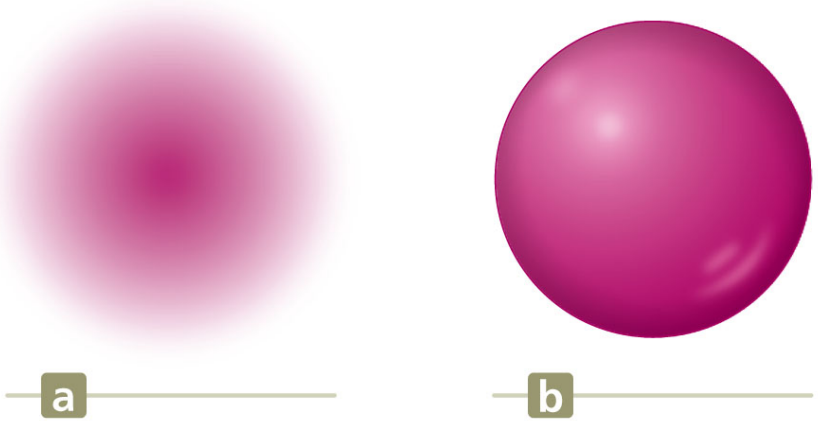
Quantum Mechanical Model (1920s): Orbitals

- Orbital: Probability map that shows <u>probability</u> of finding the electron within a certain space
- Darker areas = higher probability of finding electron
- Orbitals ≠ Orbits: Orbitals are NOT circular.
- Orbitals give NO info about electron path or how electrons move



Quantum Mechanical Model of Atom

- Orbitals do not have sharp boundaries.
- Chemists arbitrarily define an orbital's size as the area that contains 90% of the total electron probability.



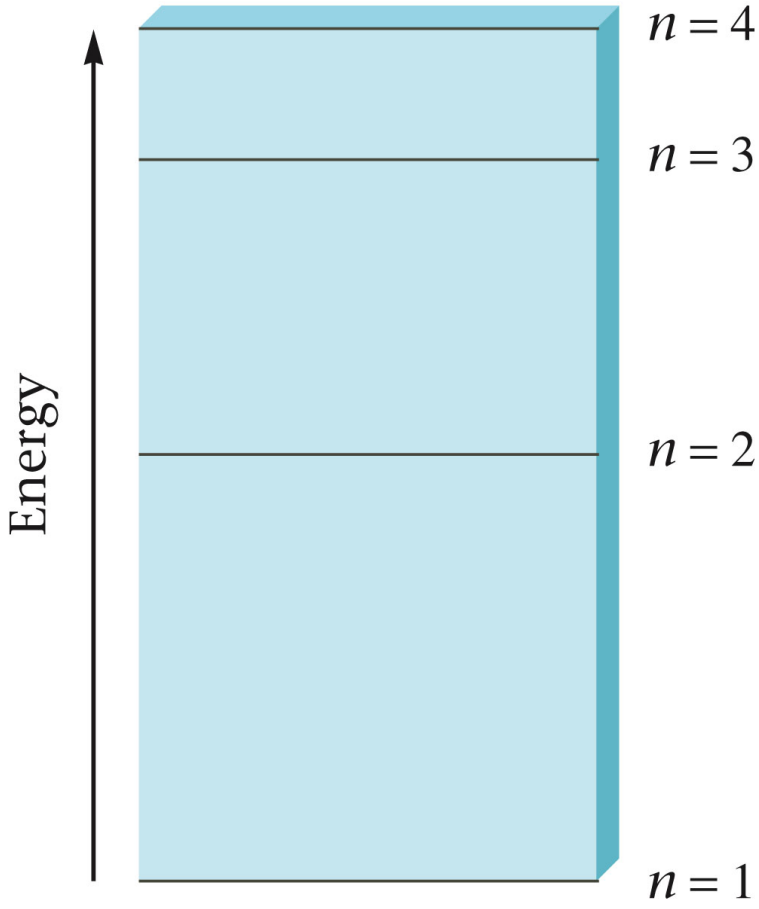
Quantum Mechanical Model of Atom

Electrons are arranged by hierarchy:

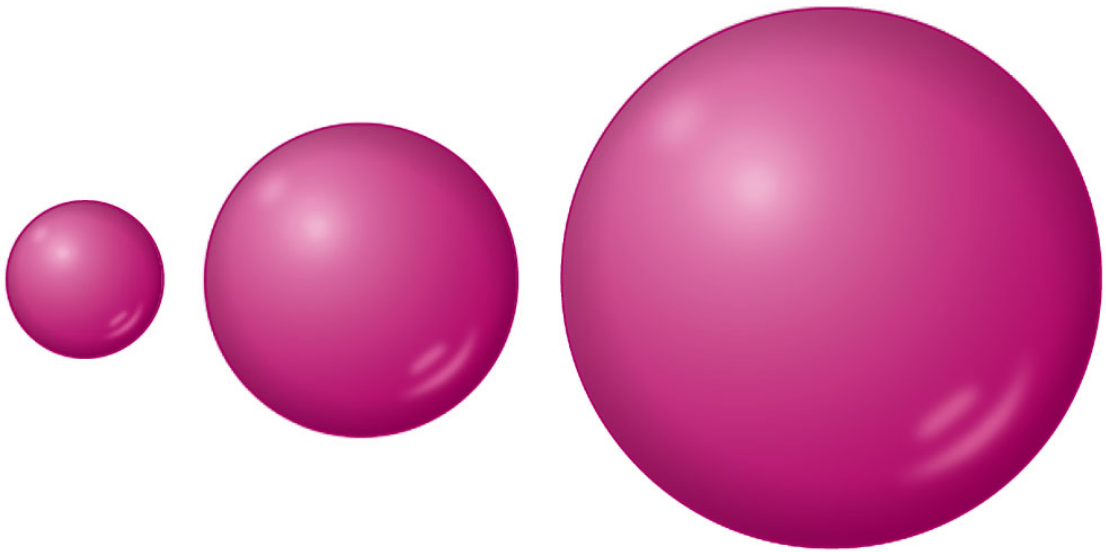
1. Principal Shells
2. Subshells
3. Orbitals
4. Electrons

Principal Shell (Shell)

- Atoms have discrete energy levels called **principal shells**:
  - Labeled with a principal quantum number *n*.
  - Higher n → higher the energy and bigger the size of principal shell (electron farther from nucleus, on average)



Shell Number Determines Energy Level and Size



n = 1 2 3

Subshell (Sublevel)

Each principal shell is composed of **subshells**.

- Number of subshells in a principal shell = principal shell number n

- Each subshell is labeled with number and letter.
  - Number indicates the principal shell that the subshell is in.
  - Letter (s, p, d, f, etc.) indicates shape of orbital(s) in the subshell.

Subshells

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

Orbitals and Their Shapes

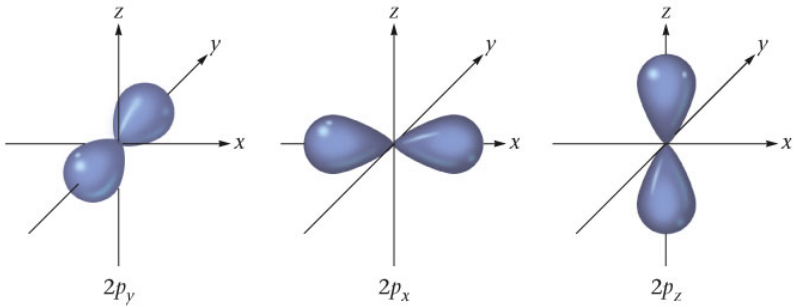
Each subshell is composed of **orbitals**.

Each s subshell is composed of one s orbital.

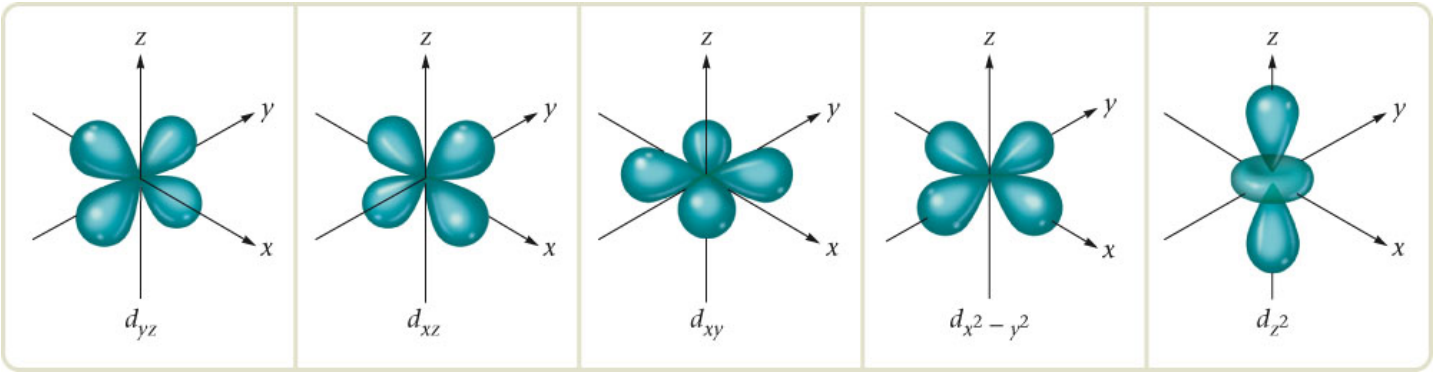
Each p subshell is composed of three p orbitals.

Each d subshell is composed of five p orbitals.

s orbitals (1)



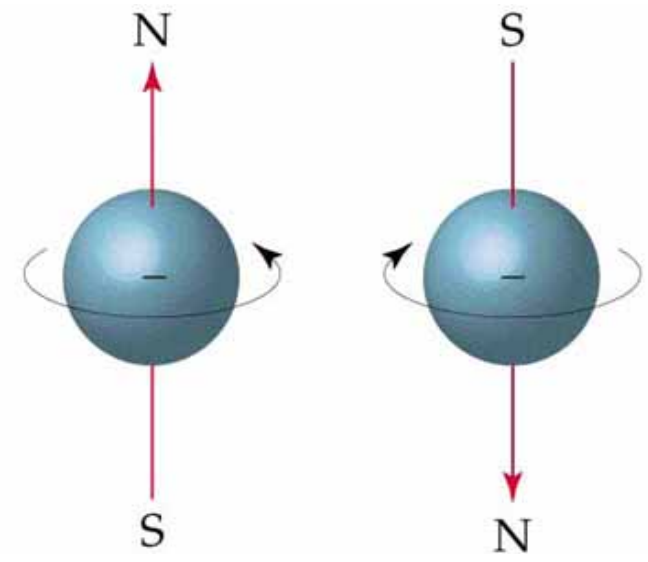
p orbitals (3)



d orbitals (5)

Electrons

- Each orbital holds a max of 2 **electrons**.
- Electrons can have two possible values of spin: up or down.



Electron Distribution in Atoms

Shell number:	1	2	3	4
Subshell designation:	s	s , p	s , p , d	s , p , d , f
Number of orbitals:	1	1 , 3	1 , 3 , 5	1 , 3 , 5 , 7
Number of electrons	2	2, 6	2, 6, 10	2, 6, 10, 14
Total electrons in shell	2	8	18	32

Questions

<u>True or False?</u>

1. An s orbital is always spherical in shape. T
2. 2s orbital is same size as 3s orbital. F
3. 3p orbitals have more lobes than 2p orbitals. F
4. Principal shell 1 has one s orbital, principal shell 2 has two s orbitals, principal shell 3 has three s orbitals, etc. F
5. The shape of a given type of orbital changes as n increases. F
6. The number of subshells = principal shell number T

Electron Configurations

Electron Configurations

Ways of Representing Electron Arrangement in Atom

1. Electron Configuration

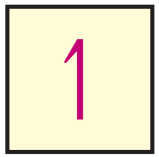
H $1s^1$

Number of electrons in subshell

Subshell

2. Orbital Diagram: Orbital is a line or box containing arrow(s) to represent electrons.

H

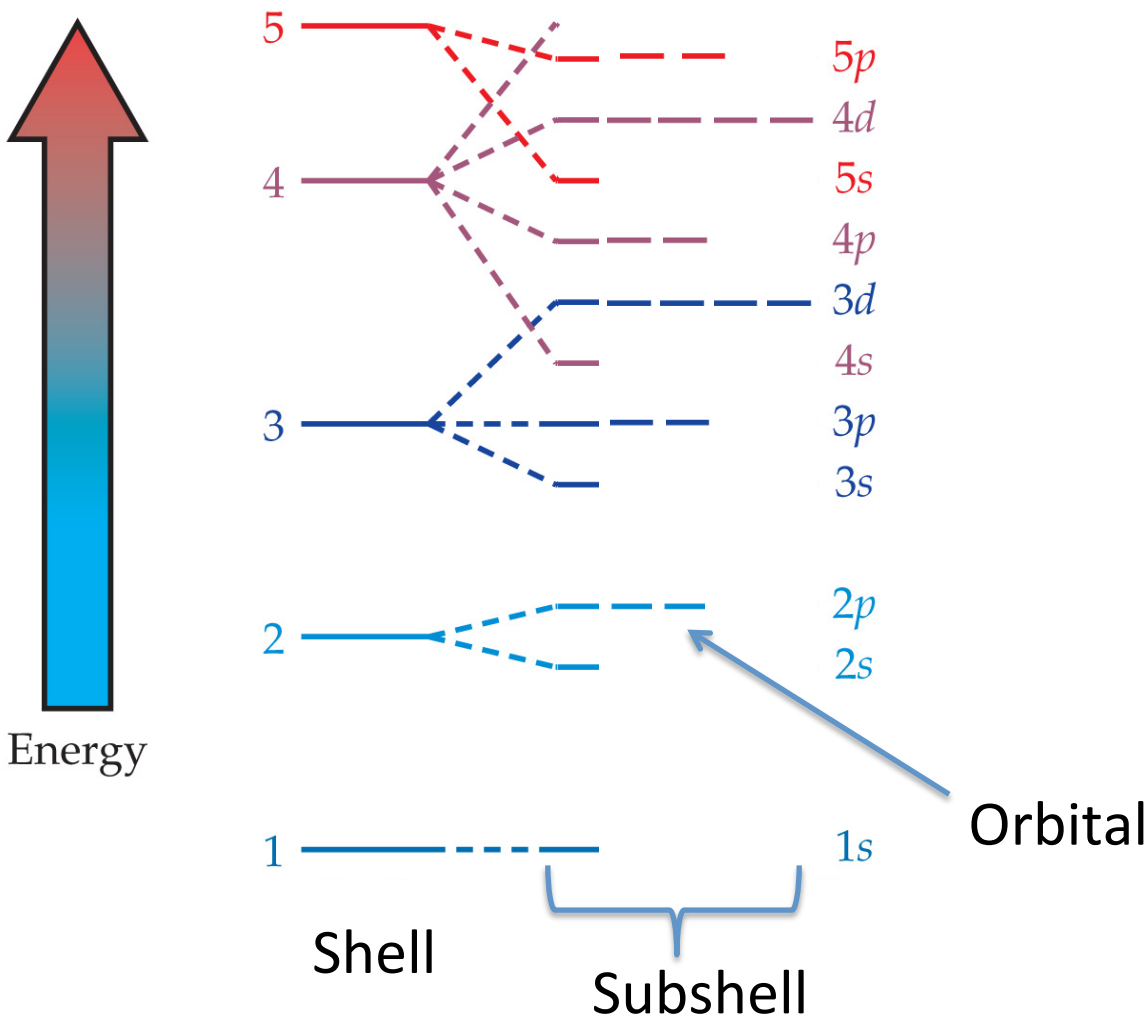


1s

Rules for Filling Orbitals with Electrons

1. Aufbau Principle: Fill lowest-energy orbitals first.
   - Shell energies: 1 < 2 < 3… (in general)
   - Subshell energies: s < p < d < f… (in general)
2. Pauli Exclusion Principle: Each orbital can hold a maximum of 2 electrons, and those 2 electrons must have opposite spins.
3. Hund's Rule: Must have maximum number of <u>unpaired</u> electrons in a set of same-energy (degenerate) orbitals.

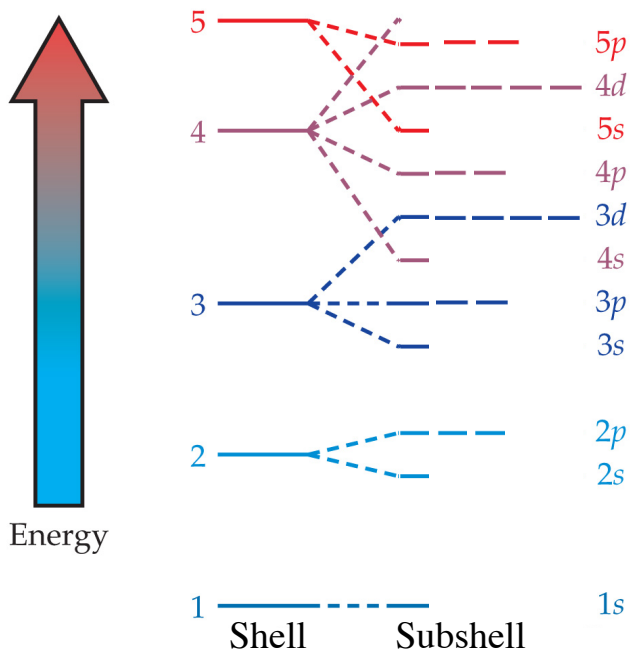
Energy Levels of Shell, Subshell, Orbital



Period 1: Electron Arrangement for H (Z = 1)

Note: Ground state configurations are shown.

	Electron configuration	Orbital diagram
H	$1s^1$	1
		1s



Period 1: Electron Arrangement for He (Z = 2)

Note: Ground state configurations are shown.

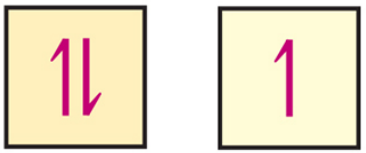
	Electron configuration	Orbital diagram
He	$1s^2$	$\boxed{\uparrow\downarrow}$ 1s

Pauli Exclusion Principle: Each orbital can hold a maximum of 2 electrons, and those 2 electrons must have opposite spins.

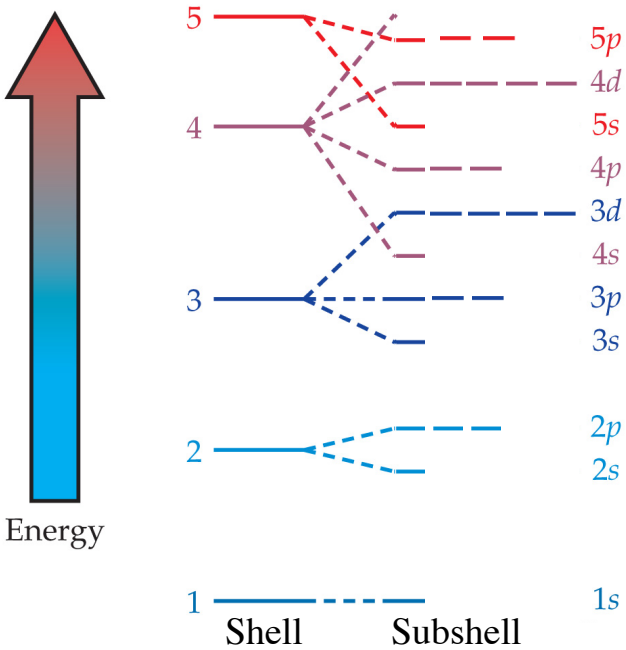
Period 2: Electron Arrangement for Li (Z=3)

Electron configuration | **Orbital diagram**

Li $1s^2 2s^1$

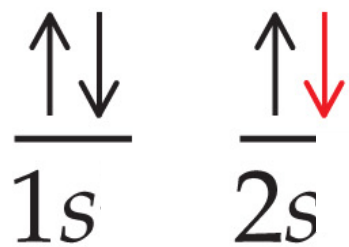


1s 2s



Electron Arrangement for Be (Z=4)

Be $1s^2\ 2s^2$

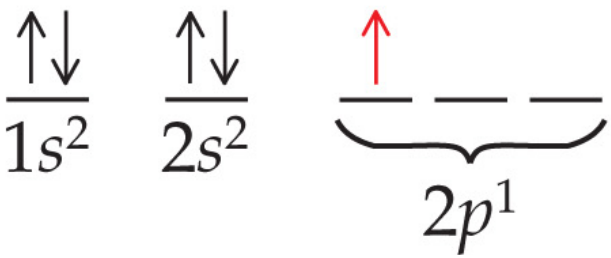


$[He]\ 2s^2$

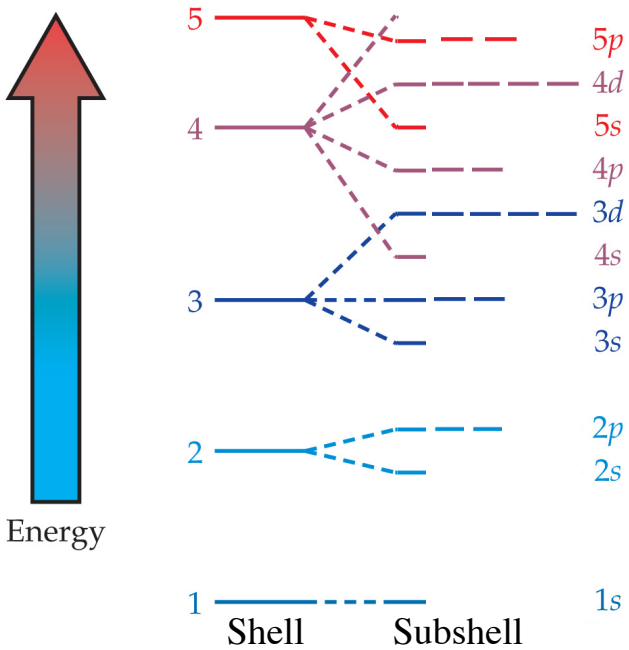
Shorthand Notation: Put in brackets last element to have completely-filled shell (noble gas)

Electron Arrangement for B (Z=5)

B $1s^2\,2s^2\,2p^1$

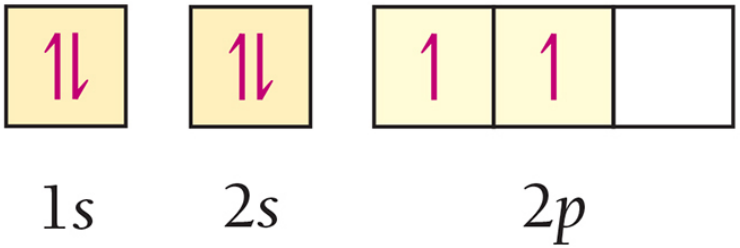


[He] $2s^2\,2p^1$



Electron Arrangement for C (Z=6)

Electron configuration		Orbital diagram
C	$1s^2 2s^2 2p^2$	

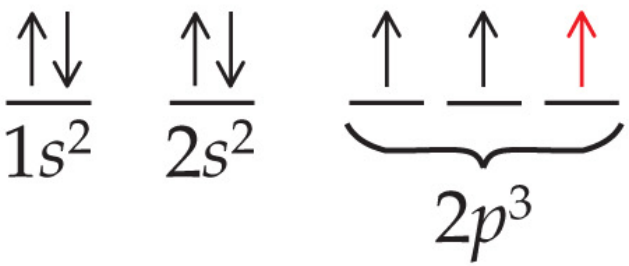


$[He]\ 2s^2\ 2p^2$

Hund's Rule: Must have maximum number of <u>unpaired</u> electrons in a set of same-energy (degenerate) orbitals.

Electron Arrangement for N (Z=7)

N $1s^2\ 2s^2\ 2p^3$



[He] $2s^2\ 2p^3$

Electron Arrangement for O (Z=8), F (Z=9), and Ne (Z=10)

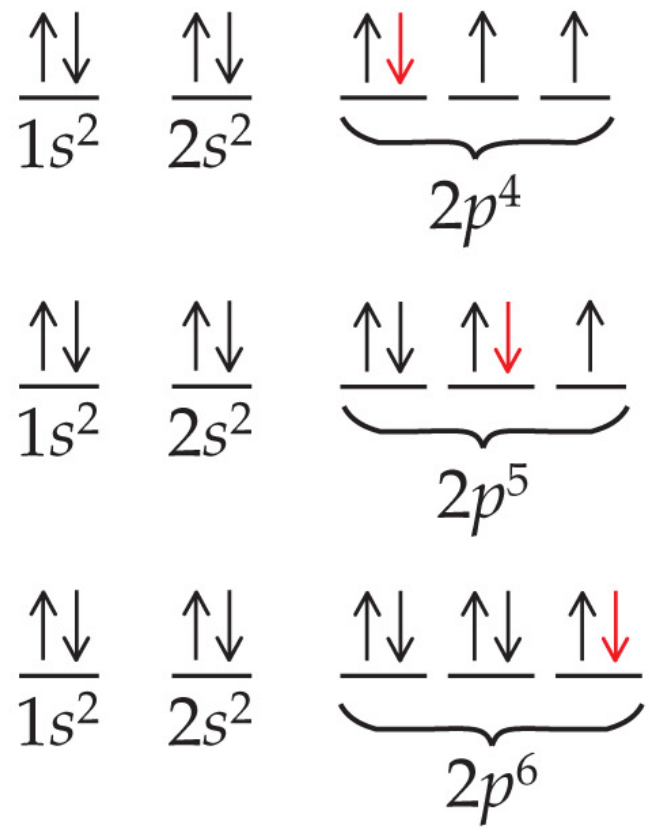
O $1s^2\,2s^2\,2p^4$

[He] $2s^2\,2p^4$

F $1s^2\,2s^2\,2p^5$

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

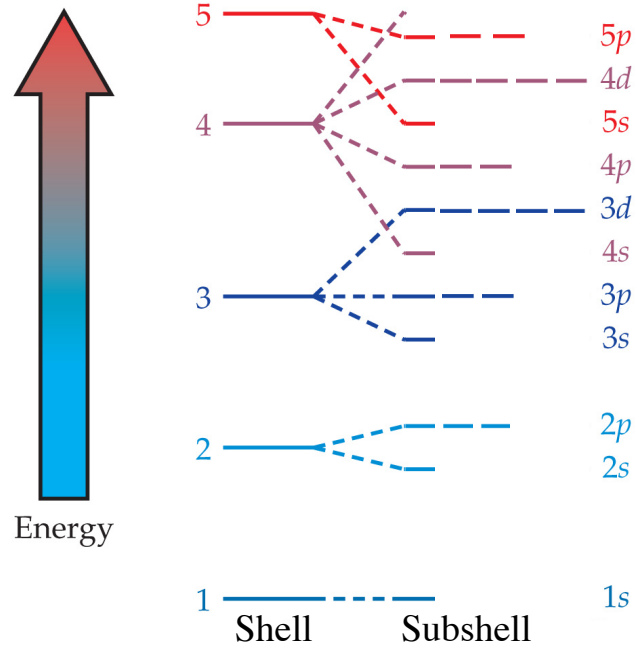
Ne $1s^2\,2s^2\,2p^6$

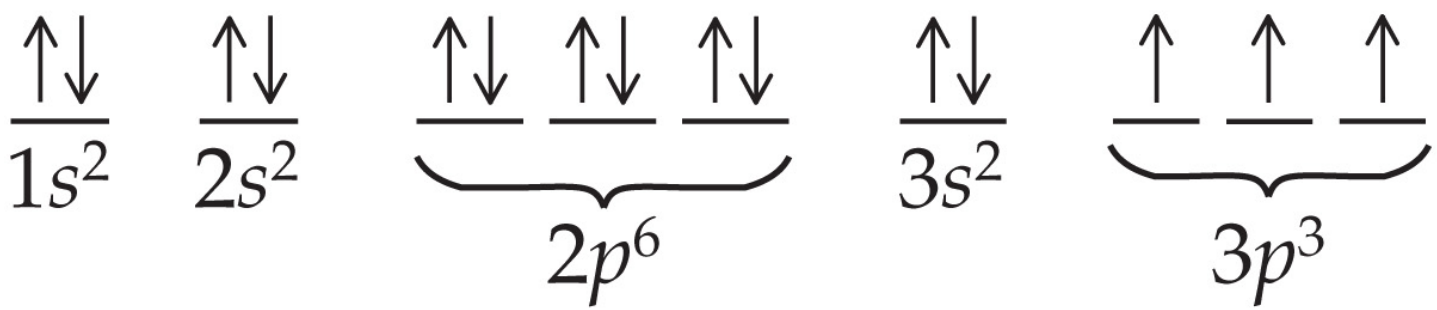


Period 3:
Electron Arrangement for P (Z =15)

P $1s^2 2s^2 2p^6 3s^2 3p^3$

[Ne] $3s^2 3p^3$





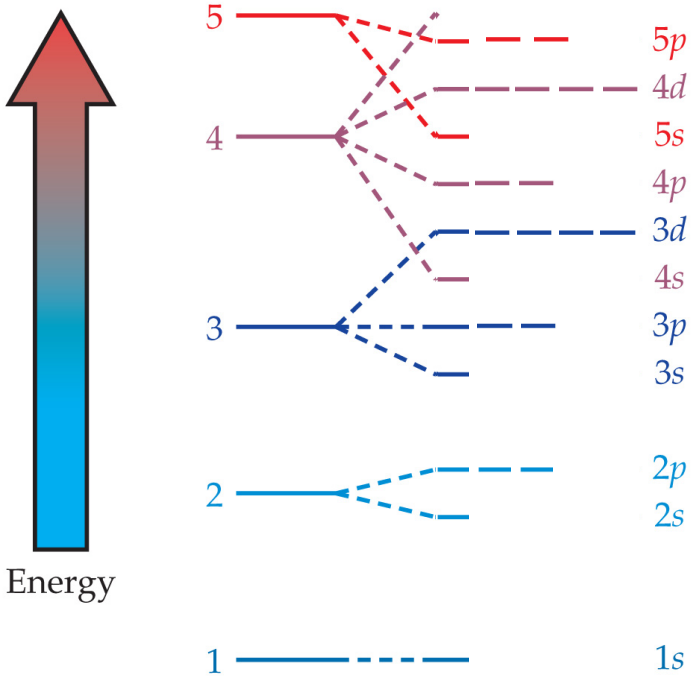
Period 4:
Electron Arrangement for K (Z=19) and Ca (Z=20)

K $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$ [Ar]$4s^1$

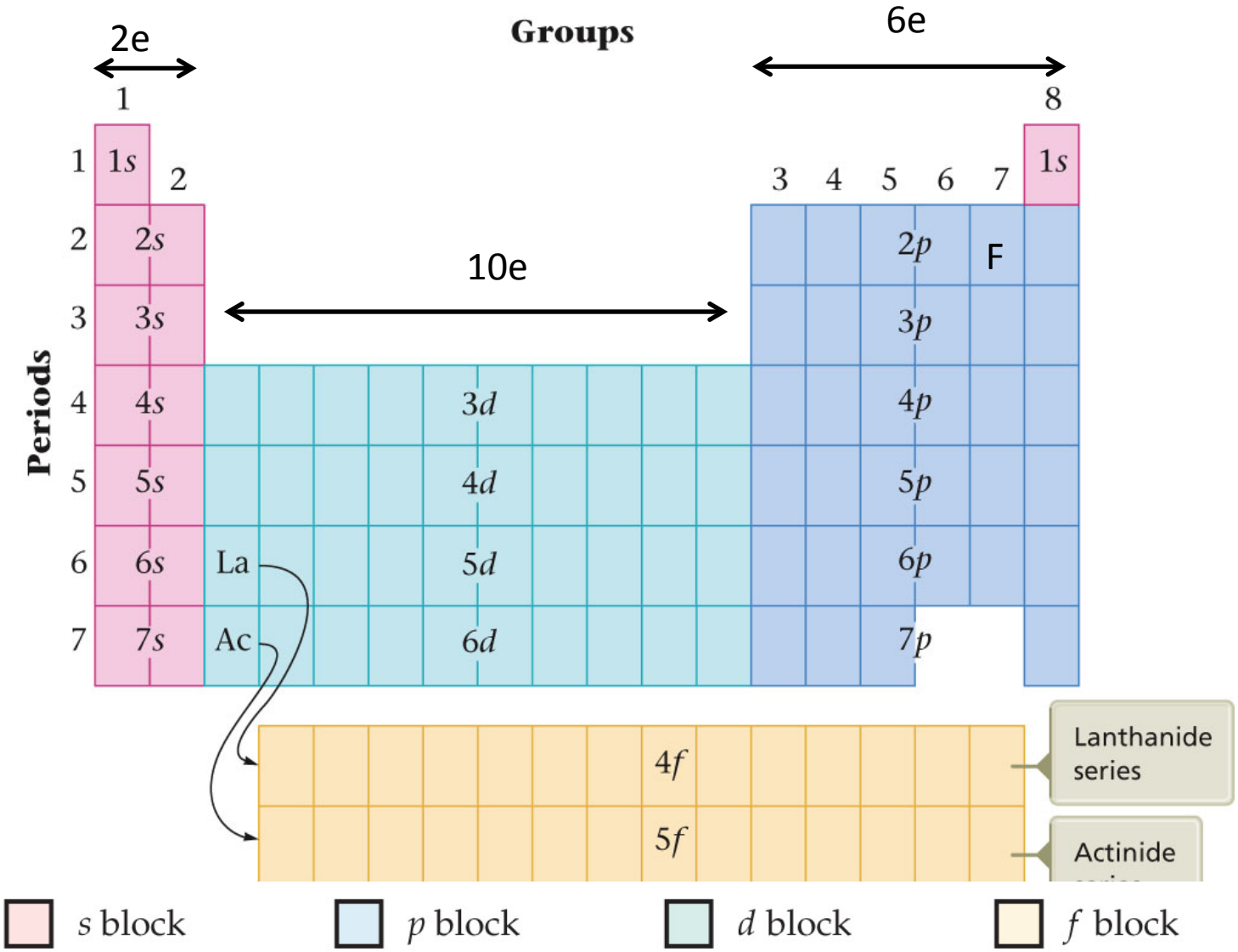
Ca $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$ [Ar]$4s^2$

Note: The ns orbitals fill before (n-1)d orbitals.

4s fills before 3d.



Periodic Table Shows Orbital Filling Order (Memorize!)



Periodic Table Shows Orbital Filling Order

- Blocks show the last subshell to be filled.
- Arrangement of blocks shows order of orbital energy levels.
- For s and p orbitals:

  Period number = n (principal shell number)

- For d orbitals:

  Period number – 1 = n

Ex Probs

Transition Metals

When starting to fill d orbitals, can have more irregularities in filling because of ns and (n-1)d energy levels are close to each other.

In first row transition metals, Cr and Cu are "irregular."

			3	4	5	6	7	8	9	10	11	12						
3	11 Na $3s^1$	12 Mg $3s^2$	3B	4B	5B	6B	7B	8B			1B	2B	13 Al $3s^23p^1$	14 Si $3s^23p^2$	15 P $3s^23p^3$	16 S $3s^23p^4$	17 Cl $3s^23p^5$	18 Ar $3s^23p^6$
4	19 K $4s^1$	20 Ca $4s^2$	21 Sc $4s^23d^1$	22 Ti $4s^23d^2$	23 V $4s^23d^3$	24 Cr $4s^13d^5$	25 Mn $4s^23d^5$	26 Fe $4s^23d^6$	27 Co $4s^23d^7$	28 Ni $4s^23d^8$	29 Cu $4s^13d^{10}$	30 Zn $4s^23d^{10}$	31 Ga $4s^24p^1$	32 Ge $4s^24p^2$	33 As $4s^24p^3$	34 Se $4s^24p^4$	35 Br $4s^24p^5$	36 Kr $4s^24p^6$
5	37 Rb $5s^1$	38 Sr $5s^2$	39 Y $5s^24d^1$	40 Zr $5s^24d^2$	41 Nb $5s^14d^4$	42 Mo $5s^14d^5$	43 Tc $5s^24d^5$	44 Ru $5s^14d^7$	45 Rh $5s^14d^8$	46 Pd $4d^{10}$	47 Ag $5s^14d^{10}$	48 Cd $5s^24d^{10}$	49 In $5s^25p^1$	50 Sn $5s^25p^2$	51 Sb $5s^25p^3$	52 Te $5s^25p^4$	53 I $5s^25p^5$	54 Xe $5s^25p^6$
6	55 Cs $6s^1$	56 Ba $6s^2$	57 La $6s^25d^1$	72 Hf $6s^25d^2$	73 Ta $6s^25d^3$	74 W $6s^25d^4$	75 Re $6s^25d^5$	76 Os $6s^25d^6$	77 Ir $6s^25d^7$	78 Pt $6s^15d^9$	79 Au $6s^15d^{10}$	80 Hg $6s^25d^{10}$	81 Tl $6s^26p^1$	82 Pb $6s^26p^2$	83 Bi $6s^26p^3$	84 Po $6s^26p^4$	85 At $6s^26p^5$	86 Rn $6s^26p^6$
7	87 Fr $7s^1$	88 Ra $7s^2$	89 Ac $7s^26d^1$	104 Rf $7s^26d^2$	105 Db $7s^26d^3$	106 Sg $7s^26d^4$	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Mc	116 Lv	117 Ts	118 Og

Valence and Core Electrons

- Valence electrons: electrons in the outermost principal shell (highest n) of an atom [for main group elements]
  - Ne: $1s^2 2s^2 2p^6$ (# of valence electrons = 8)
  - O: $1s^2 2s^2 2p^4$ (# of valence electrons = 6)
- Valence electrons are involved in the chemistry!
- For main group elements, Group no. = No. of valence electrons (except for He).
- Core electrons: electrons that are NOT in the outermost principal shell (electrons in inner shells)

Ex Probs

Electron Configurations in Last Occupied Shell

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

Elements in the same group have similar outermost shell (valence) electron configurations!

→ That's why elements of same group have similar chemical properties!

Periodic Trends

Periodic Trends

Periodic Table Shows Ion Formation Trend

- Metal elements tend to lose electrons and form cations (positive ions).

- Nonmetal elements tend to gain electrons and form anions (negative ions).

Trends in Ion Charge

Main group elements lose or gain electrons so they can form ions with noble gas electron configuration (*isoelectronic* with the nearest noble gas).

Group	Element	Elec Config	Ion Formation	Ion Elec Config	Ion Charge
1	Na	$[Ne]3s^1$	$Na \rightarrow Na^+ + e^-$	[Ne]	+1
2	Mg	$[Ne]3s^2$	$Mg \rightarrow Mg^{2+} + 2e^-$	[Ne]	+2
3	Al	$[Ne]3s^23p^1$	$Al \rightarrow Al^{3+} + 3e^-$	[Ne]	+3
6	O	$[He]2s^22p^4$	$O + 2e^- \rightarrow O^{2-}$	[Ne]	-2
7	F	$[He]2s^22p^5$	$F + 2e^- \rightarrow F^-$	[Ne]	-1

Common Ion Charges

1A	2A											3A	4A	5A	6A	7A	8A
Li^{+}														N^{3-}	O^{2-}	F^{-}	
Na^{+}	Mg^{2+}											Al^{3+}			S^{2-}	Cl^{-}	
K^{+}	Ca^{2+}				Cr^{2+} Cr^{3+}	Mn^{2+} Mn^{3+}	Fe^{2+} Fe^{3+}	Co^{2+} Co^{3+}		Cu^{+} Cu^{2+}	Zn^{2+}	Ga^{3+}				Br^{-}	
Rb^{+}	Sr^{2+}									Ag^{+}	Cd^{2+}		Sn^{2+} Sn^{4+}			I^{-}	
Cs^{+}	Ba^{2+}										Hg_2^{2+} Hg^{2+}		Pb^{2+} Pb^{4+}				

<u>Metals form cations</u>.

<u>Nonmetals form anions</u>.

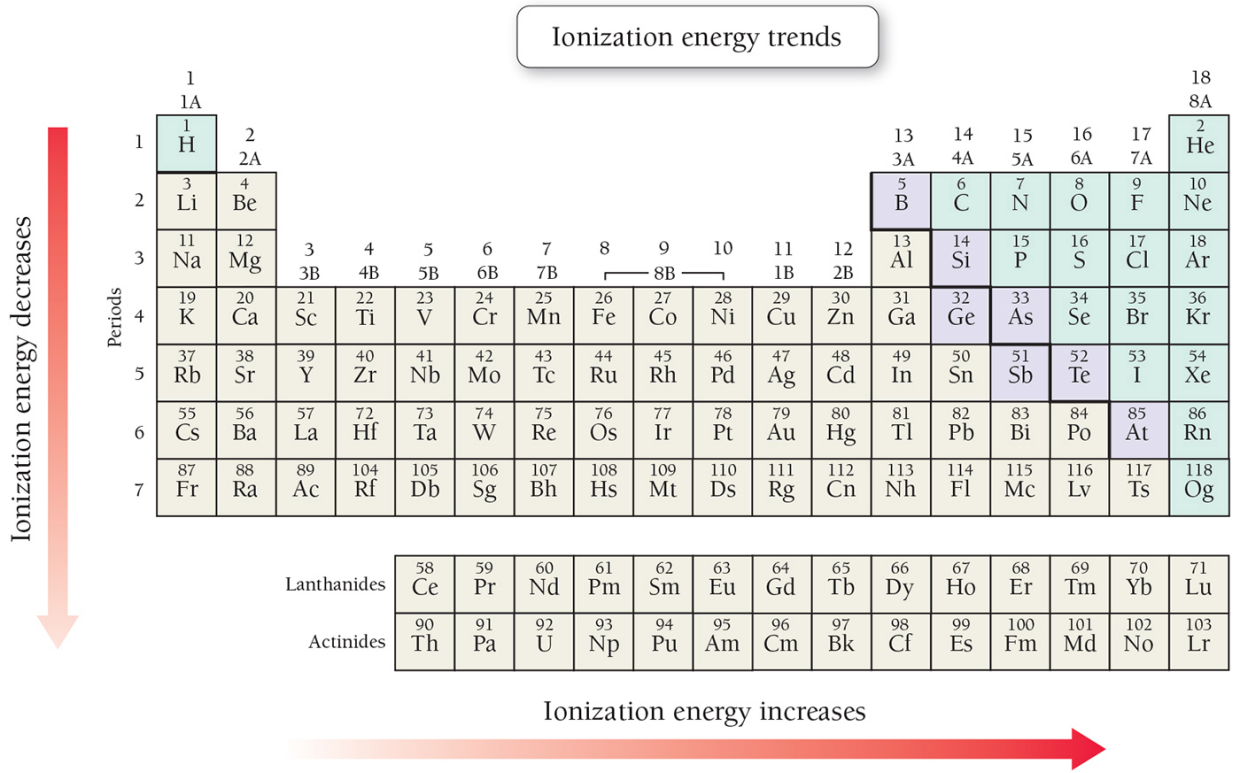
Periodic Table Shows Ionization Energy Trend

- Ionization Energy: Energy required to remove one electron from a single atom in gaseous state

$$X(g) \rightarrow X^{+}(g) + e^{-}$$ (Requires energy)

Periodic Trend in Ionization Energy

- Going across a period: Ionization energy increases → <u>Harder</u> to remove electron
- Going down a group: Ionization energy decreases → <u>Easier</u> to remove electron



Ex Probs

Trends in Metallic Character

- Metallic character: tendency to lose valence electrons (be reducing agents)

- Metals tend to have low ionization energies, so tend to lose electrons easily and form positive ions.

Periodic Table Shows Metallic Character Trend

Opposite of ionization energy trend

Metallic character trends

Least metallic

Metals
Metalloids
Nonmetals

Periods	1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8	9 8B	10	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
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	57 La	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	89 Ac	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	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	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

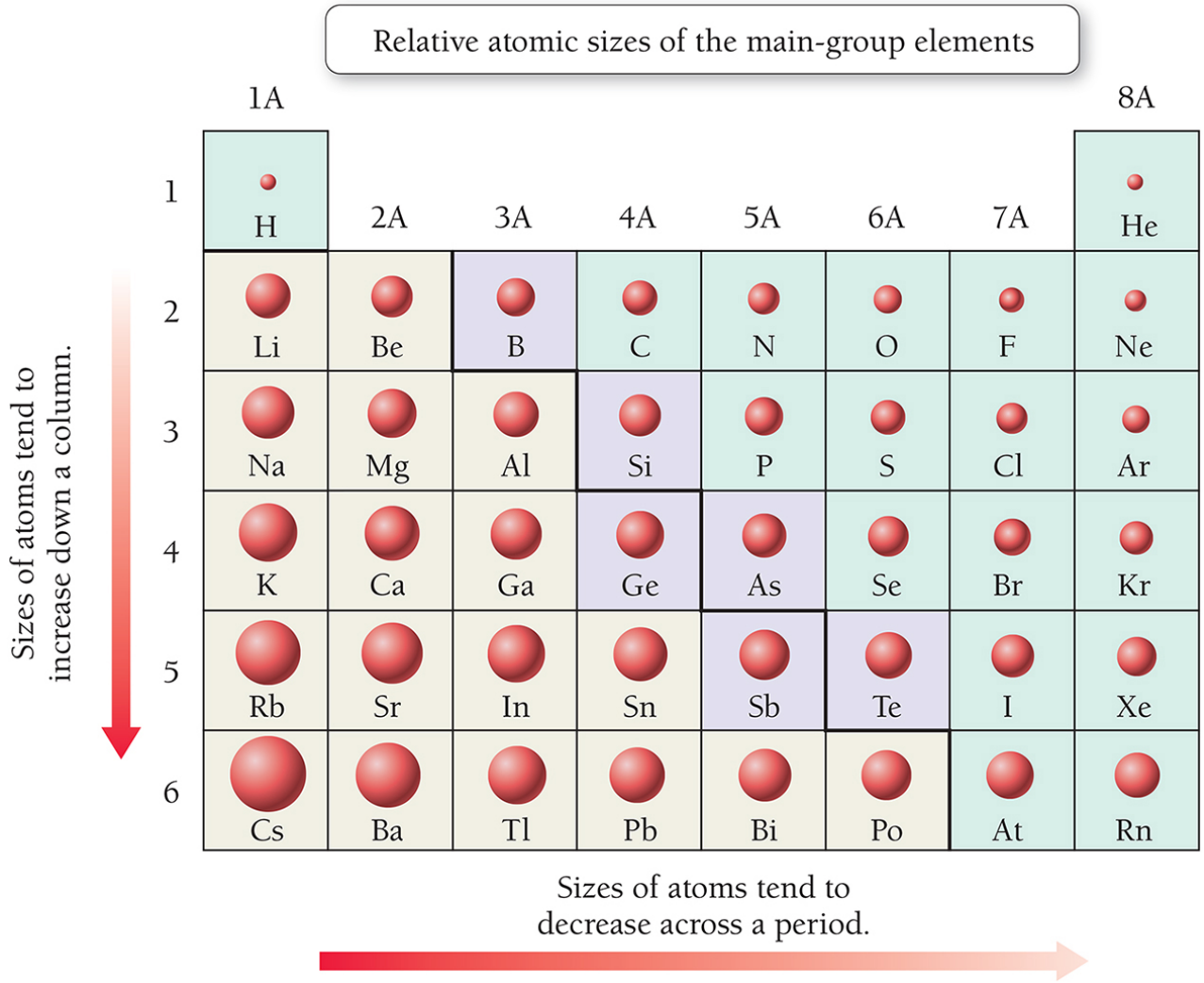
Metallic character increases

Most metallic

Metallic character decreases

Ex Probs

Periodic Table Shows Atomic Size Trend



Ex Probs