

Periodic Table of the Elements

1 H Hydrogen 1.01																	2 He Helium 4.00
3 Li Lithium 6.94	4 Be Beryllium 9.01											5 B Boron 10.81	6 C Carbon 12.01	7 N Nitrogen 14.01	8 O Oxygen 16.00	9 F Fluorine 19.00	10 Ne Neon 20.18
11 Na Sodium 22.99	12 Mg Magnesium 24.31											13 Al Aluminum 26.98	14 Si Silicon 28.09	15 P Phosphorus 30.97	16 S Sulfur 32.06	17 Cl Chlorine 35.45	18 Ar Argon 39.95
19 K Potassium 39.10	20 Ca Calcium 40.08	21 Sc Scandium 44.96	22 Ti Titanium 47.88	23 V Vanadium 50.94	24 Cr Chromium 51.99	25 Mn Manganese 54.94	26 Fe Iron 55.85	27 Co Cobalt 58.93	28 Ni Nickel 58.69	29 Cu Copper 63.55	30 Zn Zinc 65.38	31 Ga Gallium 69.72	32 Ge Germanium 72.63	33 As Arsenic 74.92	34 Se Selenium 78.97	35 Br Bromine 79.90	36 Kr Krypton 83.80
37 Rb Rubidium 85.47	38 Sr Strontium 87.62	39 Y Yttrium 88.91	40 Zr Zirconium 91.22	41 Nb Niobium 92.91	42 Mo Molybdenum 95.95	43 Tc Technetium 98.91	44 Ru Ruthenium 101.07	45 Rh Rhodium 102.91	46 Pd Palladium 106.42	47 Ag Silver 107.87	48 Cd Cadmium 112.41	49 In Indium 114.82	50 Sn Tin 118.71	51 Sb Antimony 121.76	52 Te Tellurium 127.6	53 I Iodine 126.90	54 Xe Xenon 131.29
55 Cs Cesium 132.91	56 Ba Barium 137.33	57-71 Lanthanides	72 Hf Hafnium 178.49	73 Ta Tantalum 180.95	74 W Tungsten 183.85	75 Re Rhenium 186.21	76 Os Osmium 190.23	77 Ir Iridium 192.22	78 Pt Platinum 195.08	79 Au Gold 196.97	80 Hg Mercury 200.59	81 Tl Thallium 204.38	82 Pb Lead 207.20	83 Bi Bismuth 208.98	84 Po Polonium [208.98]	85 At Astatine 209.98	86 Rn Radon 222.02
87 Fr Francium 223.02	88 Ra Radium 226.03	89-103 Actinides	104 Rf Rutherfordium [261]	105 Db Dubnium [262]	106 Sg Seaborgium [266]	107 Bh Bohrium [264]	108 Hs Hassium [277]	109 Mt Meitnerium [278]	110 Ds Darmstadtium [281]	111 Rg Roentgenium [282]	112 Cn Copernicium [285]	113 Nh Nihonium [286]	114 Fl Flerovium [289]	115 Mc Moscovium [289]	116 Lv Livermorium [293]	117 Ts Tennessine [294]	118 Og Oganesson [294]

General Chemistry

202-SN1-RE

with Olivia Bibollet-Bahena

Office: 5th floor

57 La Lanthanum 138.91	58 Ce Cerium 140.12	59 Pr Praseodymium 140.91	60 Nd Neodymium 144.24	61 Pm Promethium 144.91	62 Sm Samarium 150.36	63 Eu Europium 151.96	64 Gd Gadolinium 157.25	65 Tb Terbium 158.93	66 Dy Dysprosium 162.50	67 Ho Holmium 164.93	68 Er Erbium 167.26	69 Tm Thulium 168.93	70 Yb Ytterbium 173.06	71 Lu Lutetium 174.97
89 Ac Actinium 227.03	90 Th Thorium 232.04	91 Pa Protactinium 231.04	92 U Uranium 238.03	93 Np Neptunium 237.05	94 Pu Plutonium 244.06	95 Am Americium 243.06	96 Cm Curium 247.07	97 Bk Berkelium 247.07	98 Cf Californium 251.08	99 Es Einsteinium [254]	100 Fm Fermium 257.10	101 Md Mendelevium 258.10	102 No Nobelium 259.10	103 Lr Lawrencium [262]

- Alkali Metal
- Alkaline Earth
- Transition Metal
- Basic Metal
- Metalloid
- Nonmetal
- Halogen
- Noble Gas
- Lanthanide
- Actinide

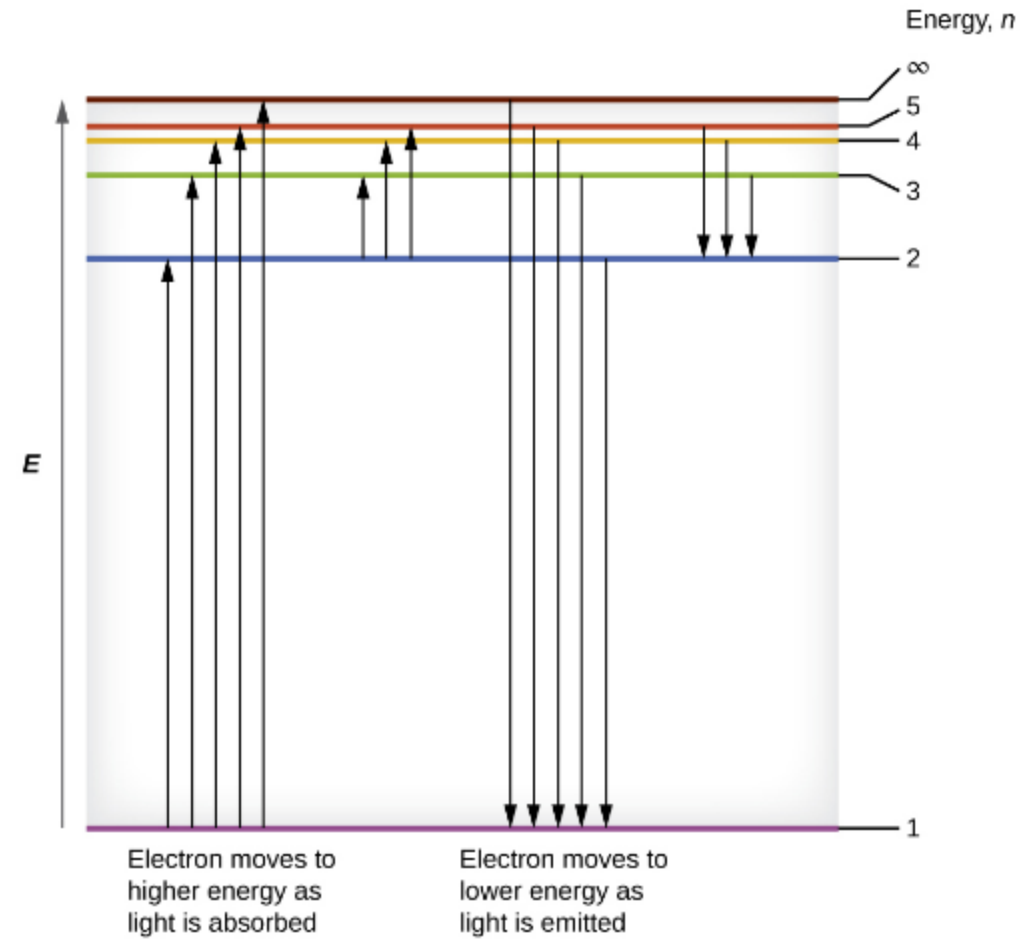


Figure 6.15 The horizontal lines show the relative energy of orbits in the Bohr model of the hydrogen atom, and the vertical arrows depict the energy of photons absorbed (left) or emitted (right) as electrons move between these orbits.

Standing Waves and the Wave Equation

- By 1920's, the Bohr model was deemed lacking
- de Broglie's addition to model was intriguing:
 - e- have wave properties
- Led to Schrodinger and Heisenberg's critical work
 - They focused on the wave properties of the e-

Heisenberg Uncertainty Principle



Werner Karl Heisenberg
(1901 –1976)
German theoretical physicist

There is a fundamental limitation to how precisely we can know both the position and momentum of a particle at a given time

We can only predict the probability of where we are going to find the electron!

Quantum Mechanical Model of the Atom

- We do not know the detailed pathway of an electron.
- Heisenberg **uncertainty principle**:
 - There is a fundamental limitation to just how precisely we can know both the position and momentum of a particle at a given time.

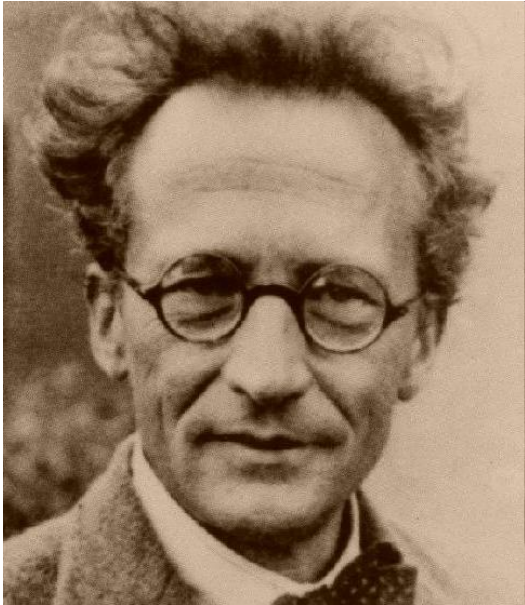
$$\Delta x \cdot \Delta(mv) \geq \frac{h}{4\pi}$$

Δx = uncertainty in a particle's position

$\Delta(mv)$ = uncertainty in a particle's momentum

h = Planck's constant

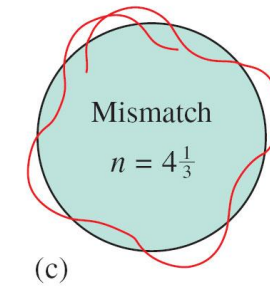
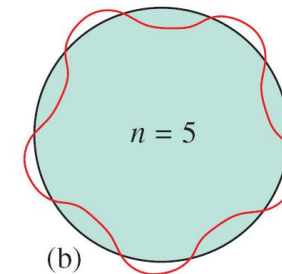
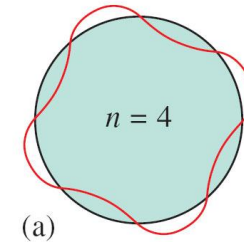
Standing Waves and the Wave Equation



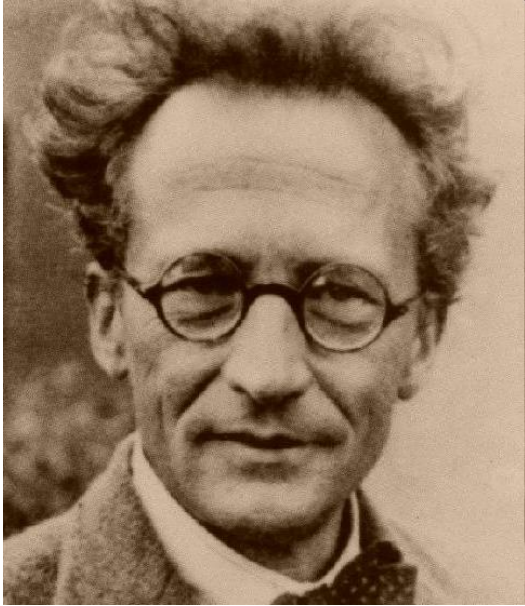
Erwin Schrodinger

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

Schrodinger developed a wave equation in analogy to the classical wave equation!



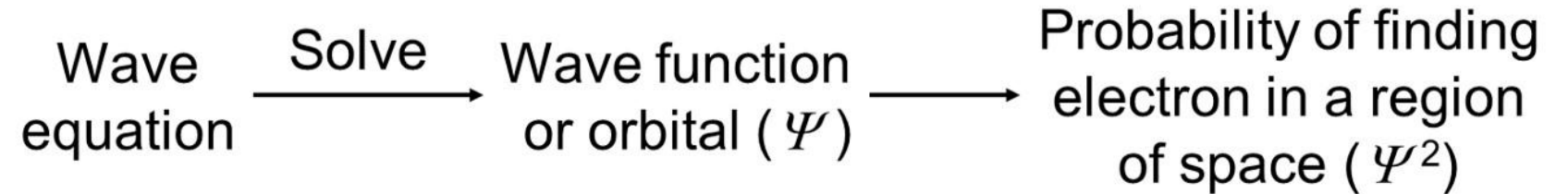
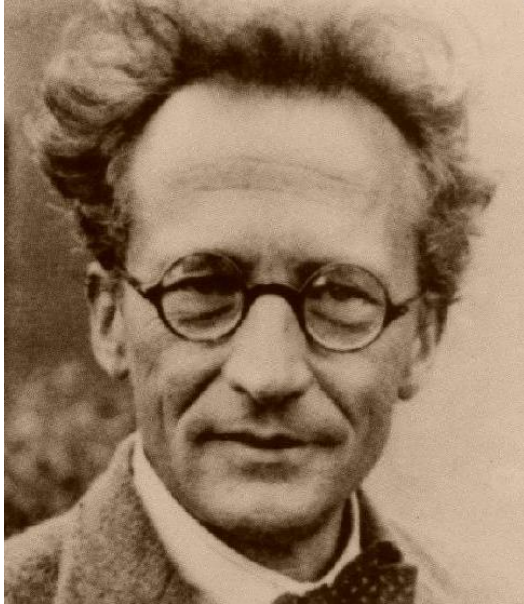
Standing Waves and the Wave Equation



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

Standing Waves and the Wave Equation

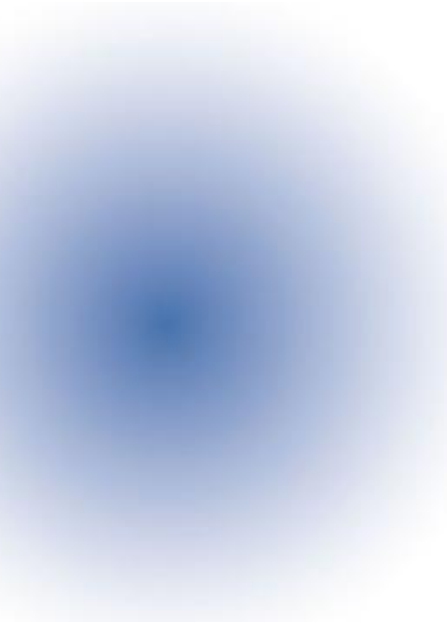


Schrodinger

The square of the wave function describes the **probability** of finding an electron at a particular point in space!

Orbitals and Electron Probability: Its Physical Meaning

$$\psi^2$$



- The square of the function indicates the probability of finding an electron near a particular point in space.
- Often visualized using a probability distribution plot – intensity of color is used to indicate the probability value near a given point in space.

Atomic orbitals represent the three-dimensional area in space that **encloses 90%** of the total electron probability

Orbitals

Trajectoire
classique



Diagramme de distribution
des probabilités de la mécanique quantique



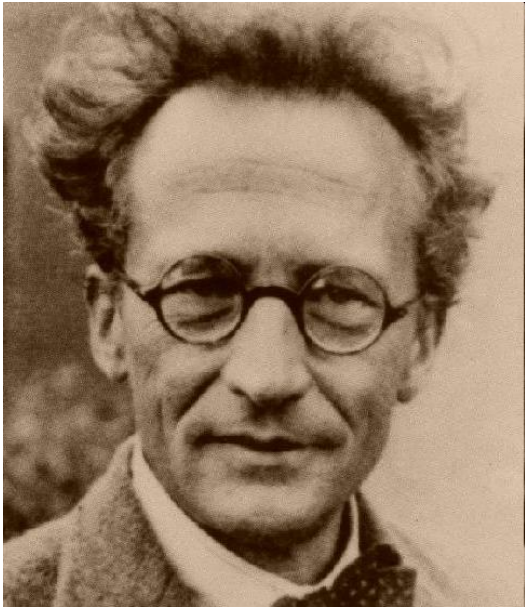
Relative Orbital Size

- An orbital (and its size) is difficult to define precisely.
- An orbital is a wave function.
- Picture an orbital as a three-dimensional electron density map.
- Hydrogen 1s orbital:
 - Radius of the sphere encloses 90% of the total electron probability.
 - We don't know the exact path the electron takes within this sphere.

Quantum Numbers

Electrons and Quantum Numbers

- To describe an electron, we use **four** quantum numbers:



1. Principal quantum number (n)
2. Angular momentum quantum number (l)
3. Magnetic quantum number (m_l)
4. Spin magnetic quantum number (m_s)

The first three quantum numbers describe the orbital.

Orbitals and Quantum Numbers

Principal quantum number, n , takes integer values: 1, 2, 3, ...

- n is related to the size and energy of the orbital.
- As n increases, the orbital becomes larger.
- An increase in n means higher energy because electron is less tightly bound to the nucleus.

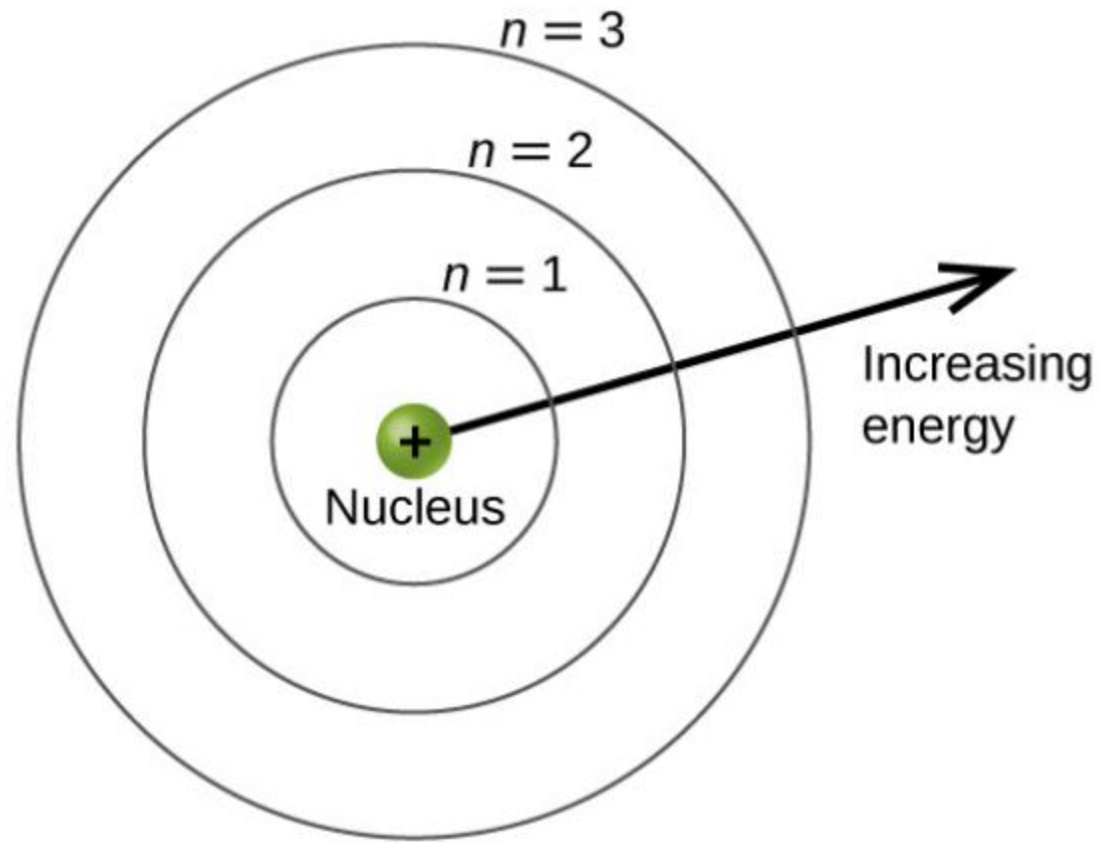


FIGURE 6.19 Different shells are numbered by principal quantum numbers.

Orbitals and Quantum Numbers

Angular momentum quantum number, l , takes integer values 0 to $n - 1$ for each value of n

- l is related to the shape of the atomic orbital
- Values of l are commonly assigned letters

Value of ℓ	0	1	2	3	4
Letter Used	<i>s</i>	<i>p</i>	<i>d</i>	<i>f</i>	<i>g</i>

Mnemonic : Some Poor Dumb Fool

Orbitals and Quantum Numbers

Magnetic quantum number, m_l , takes values between l and $-l$ (including zero)

- m_l is related to the orientations of the orbital in space

Orbitals and Quantum Numbers

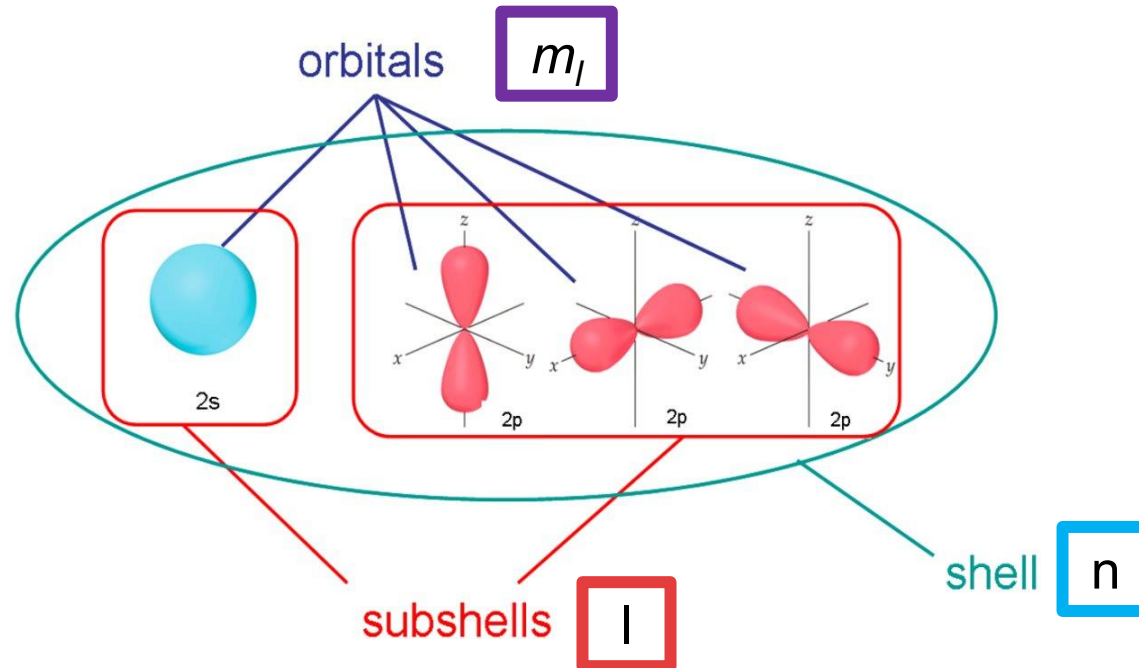
The 1s orbital of hydrogen is the lowest energy orbital, but there are others.

n	ℓ	Orbital Designation	m_ℓ	Number of Orbitals
1	0	1s	0	1
2	0	2s	0	1
	1	2p	-1, 0, +1	3
3	0	3s	0	1
	1	3p	-1, 0, 1	3
	2	3d	-2, -1, 0, 1, 2	5
4	0	4s	0	1
	1	4p	-1, 0, 1	3
	2	4d	-2, -1, 0, 1, 2	5
	3	4f	-3, -2, -1, 0, 1, 2, 3	7

Quantum numbers and orbitals

Quantum numbers : n , l et m_l .

These 3 numbers characterise a volume in space where there is a probability of finding an electron (i.e. an orbital).



The total number of orbitals at a certain energy level is equal to n^2 .

Exercise

For principal quantum level $n = 3$, determine the number of allowed subshells (different values of l), and give the designation of each.

Exercise

Is the set of quantum numbers shown below valid?

$$n = 5, \quad l = 3, \quad m_l = 4$$

Example

Keeping in mind the rules that quantum numbers must follow, can an electron be described by the following sets of quantum numbers? Why or why not?

Rules:

$$n = 1, 2, 3\dots$$

$$l = 0, 1, 2 \dots (n - 1)$$

$$m_l = 0, \pm 1, \pm 2, \pm l$$

$$n = 2, l = 1, m = -1$$

$$n = 1, l = 1, m = +1$$

$$n = 3, l = 1, m = -3$$

Exercise

Keeping in mind the rules that quantum numbers must follow, can an electron be described by the following sets of quantum numbers? Why or why not?

Rules:

$$n = 1, 2, 3 \dots$$

$$l = 0, 1, 2 \dots (n - 1)$$

$$m_l = 0, \pm 1, \pm 2, \pm l$$

$$n = 2, l = 1, m = -2$$

$$n = 3, l = 2, m = +2$$

$$n = 4, l = 3, m = +3$$

$$n = 5, l = 2, m = +3$$

Exercise

Complete the following sets of quantum numbers:

Hint: multiple answers are possible

Rules:

$$n = 1, 2, 3 \dots$$

$$l = 0, 1, 2 \dots (n - 1)$$

$$m_l = 0, \pm 1, \pm 2, \pm l$$

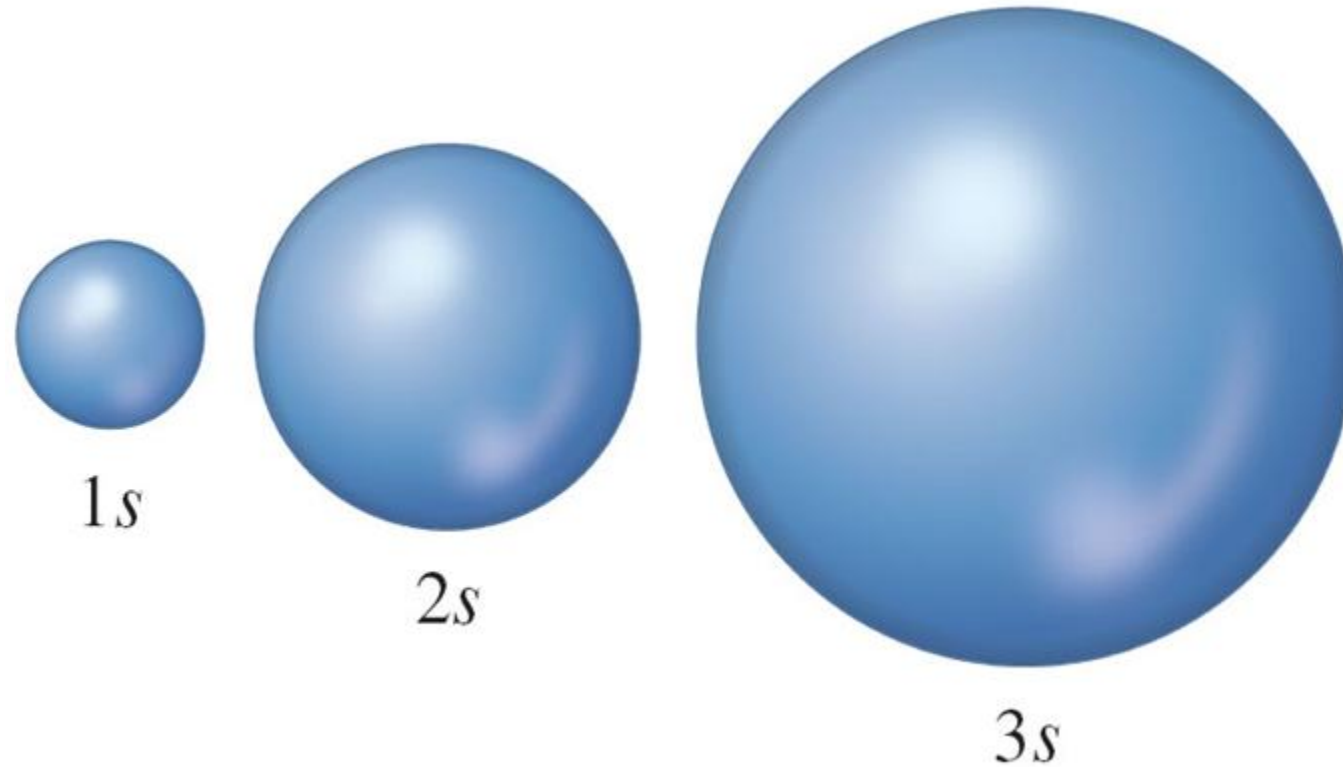
$$n = 3, l = 1, m = ?$$

$$n = 4, l = ?, m = -2$$

$$n = ?, l = 3, m = ?$$

Orbital Shapes and Energies

Orbitals Shapes and Energies



The **s orbital** is a spherically symmetrical distribution of electron probability.

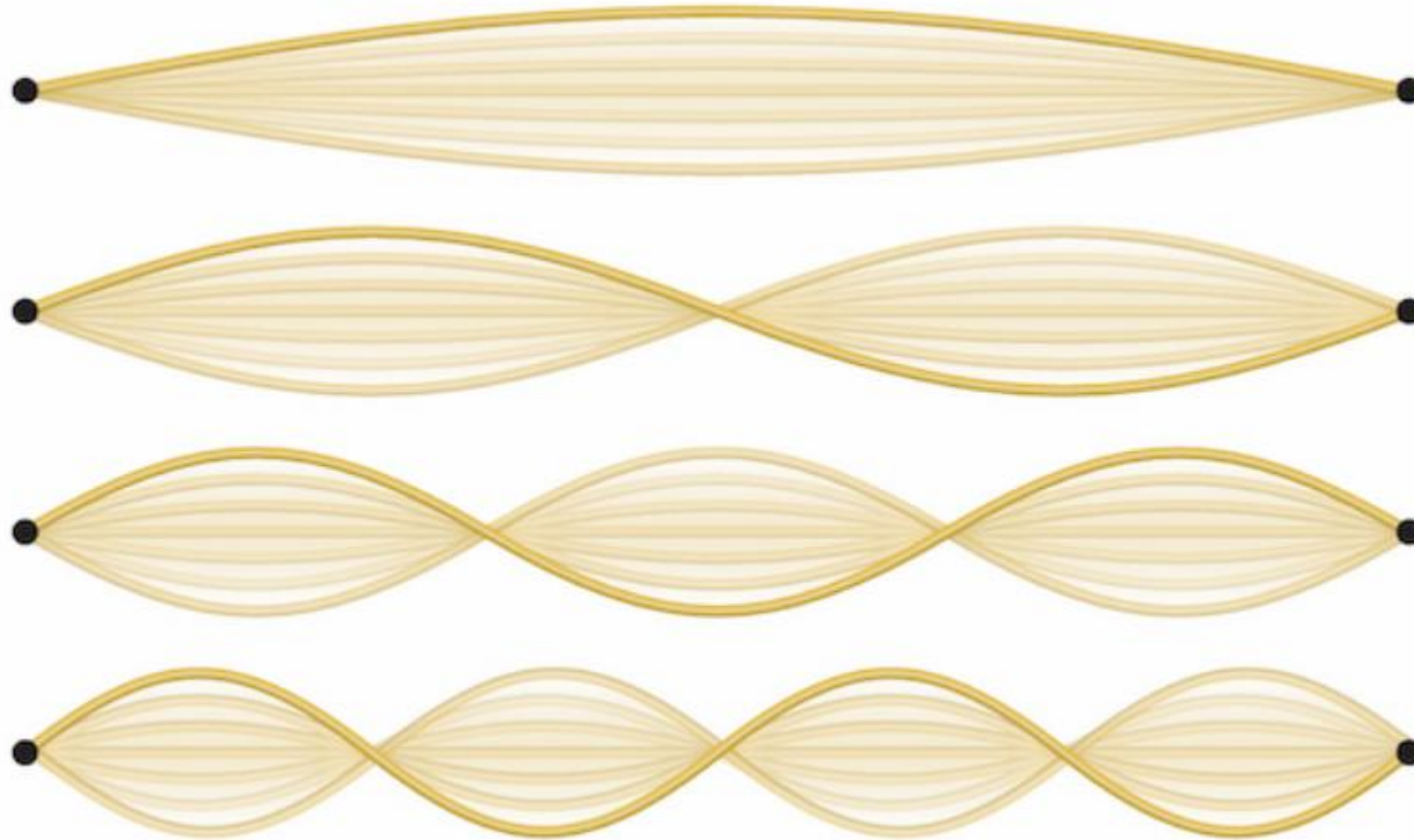
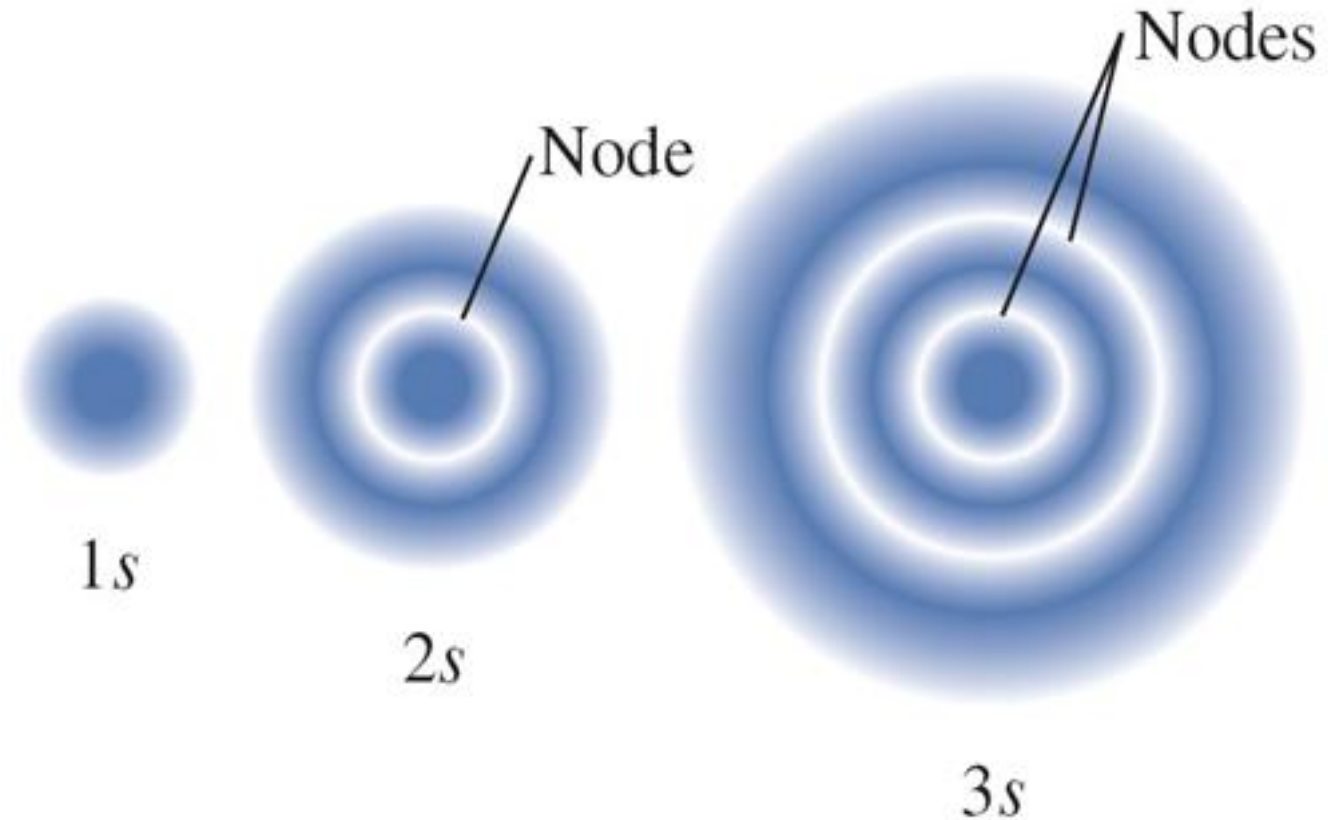


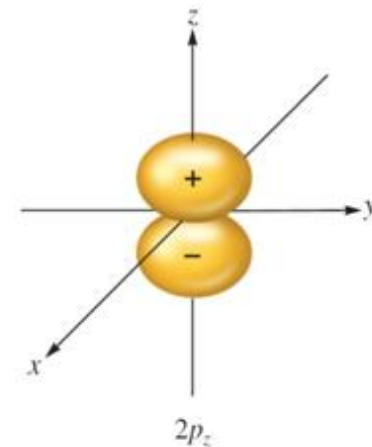
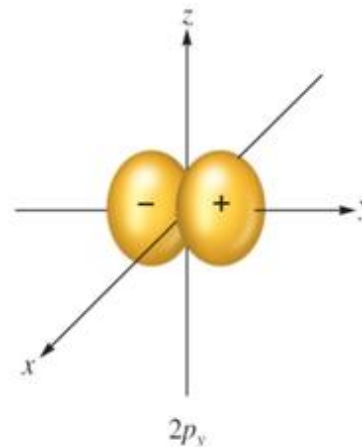
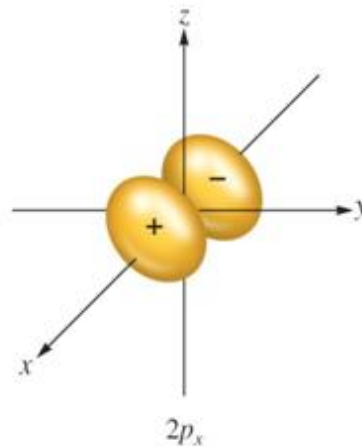
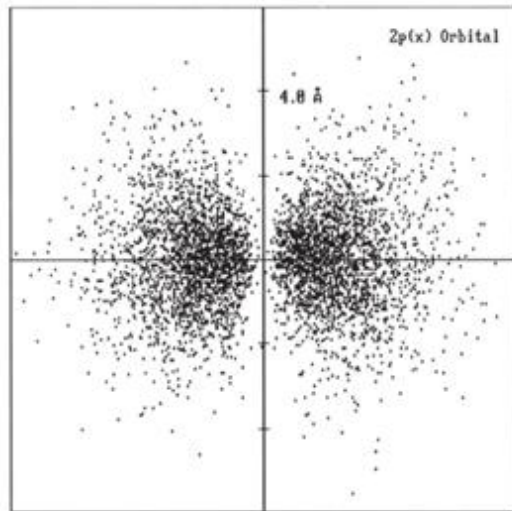
Figure 6.7 A vibrating string shows some one-dimensional standing waves. Since the two end points of the string are held fixed, only waves having an integer number of half-wavelengths can form. The points on the string between the end points that are not moving are called the nodes.

Orbitals Shapes and Energies



Orbitals can have nodes – zero probability of finding an electron (also called nodal surfaces).

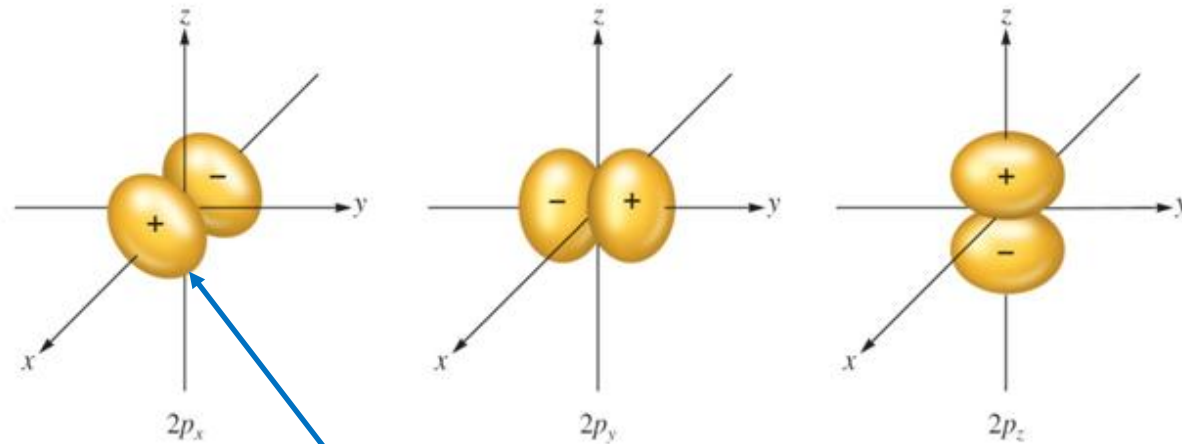
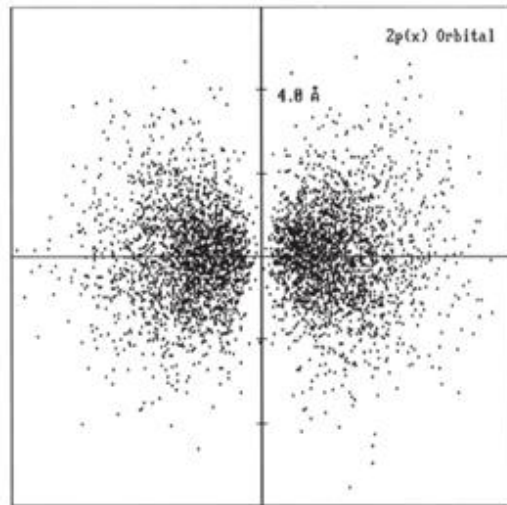
Orbitals Shapes and Energies



The ***p orbitals*** are two lobes separated by a node (called a nodal plane) at the nucleus. There are 3 orientations in space. How are they represented?

p orbitals start at $n = 2$.

Orbitals Shapes and Energies

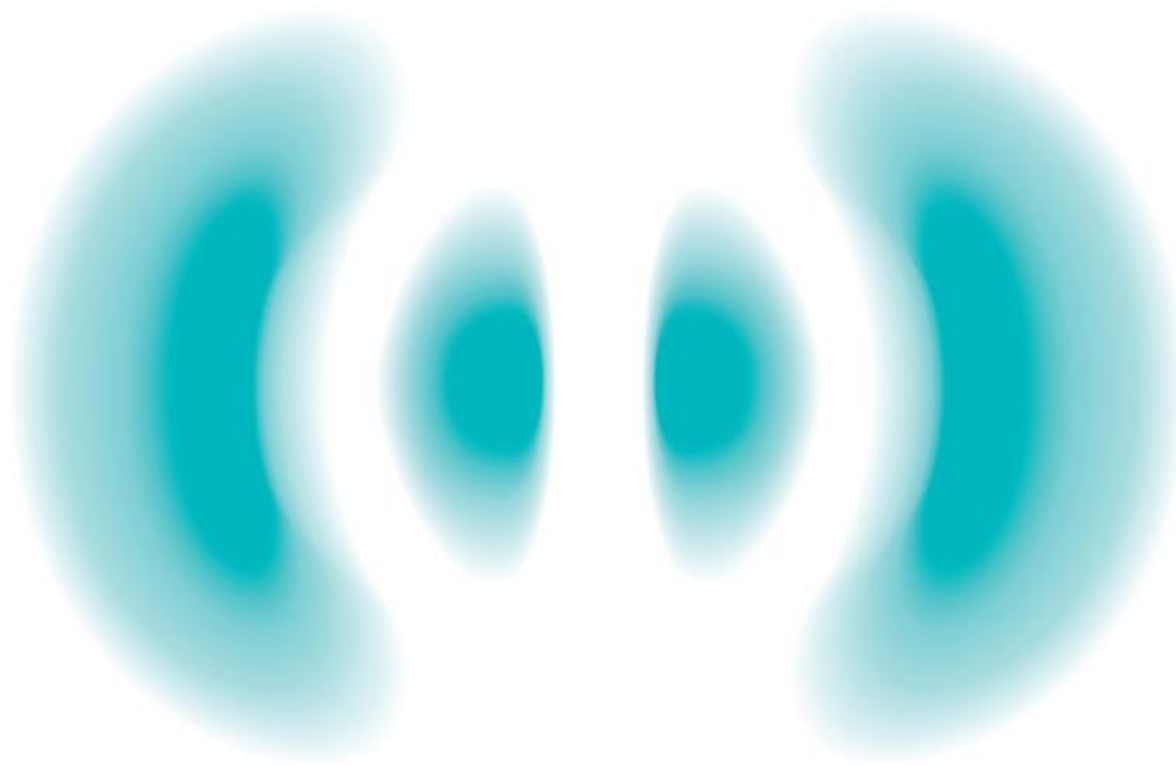


Sign of the wave function!

Signs are **not charges**; they are mathematical signs of the wave functions.

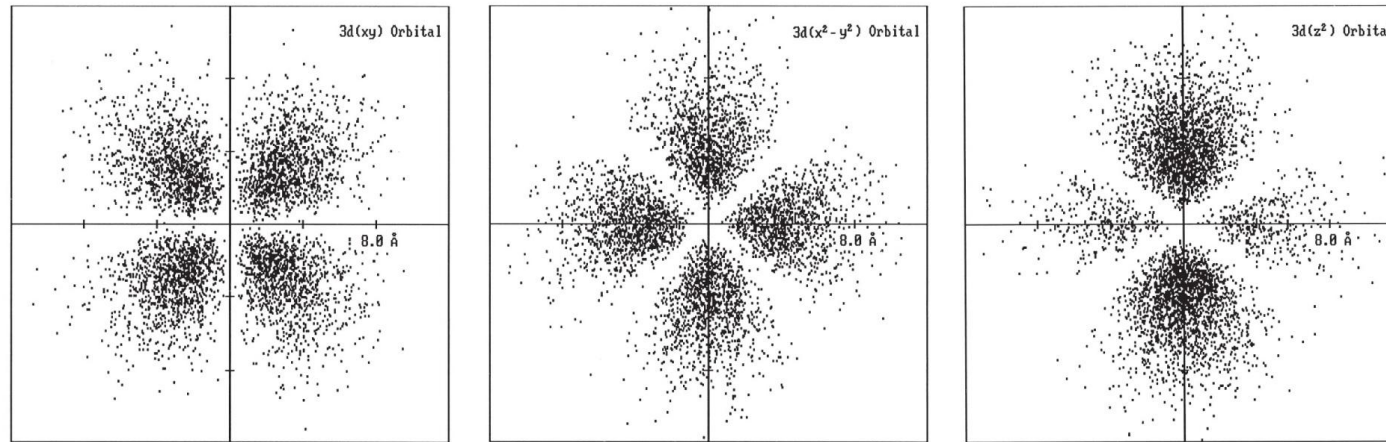
Orbitals Shapes and Energies

p orbitals with higher principal quantum numbers also have internal nodes.

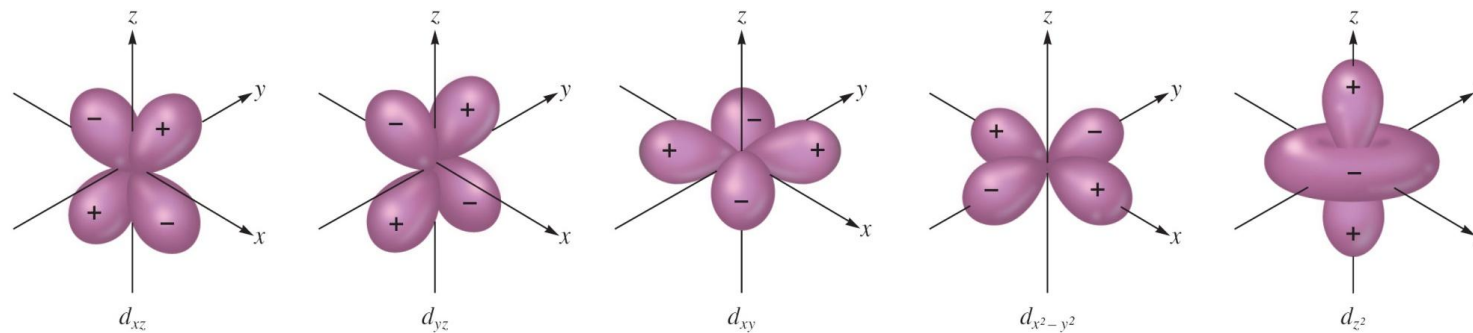


Cross section of the electron probability distribution for a $3p$ orbital shown.

Orbitals Shapes and Energies



(a)

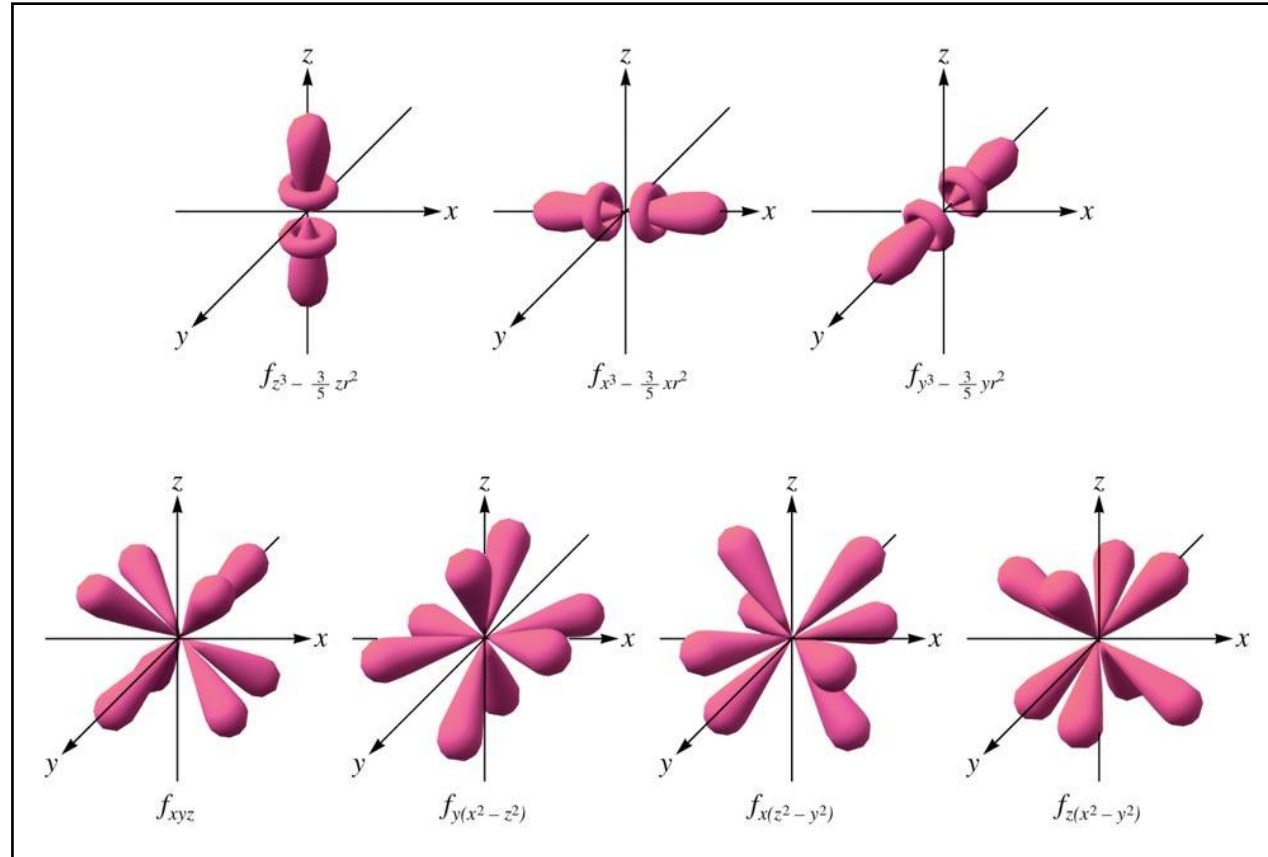


(b)

The shapes of the ***d orbitals*** get more complex and include 5 orientations in space.

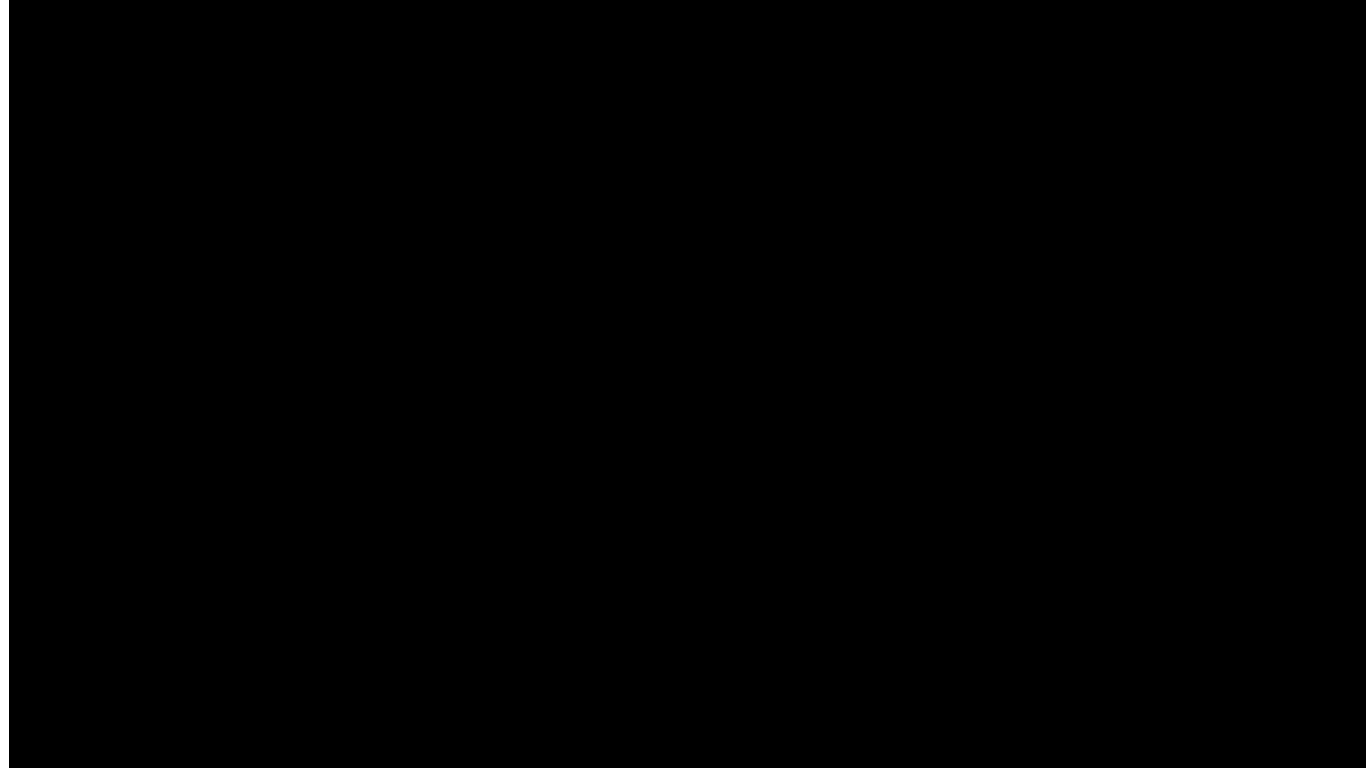
d orbitals start at $n = 3$

Representation of the 4*f* Orbitals in Terms of Their Boundary Surfaces



The shapes of the *f* orbitals are even more complex and include 7 orientations in space.
f orbitals start at $n = 4$

Orbitals



Video found at: <https://youtu.be/K-jNgq16jEY?si=SKIzswKIFVEEapVf>

The Shapes of Atoms

- If the orbitals are all different shaped, and if most of the volume of an atom is empty space diffusely occupied by electrons in these orbitals, then why do we often depict atoms as spheres?

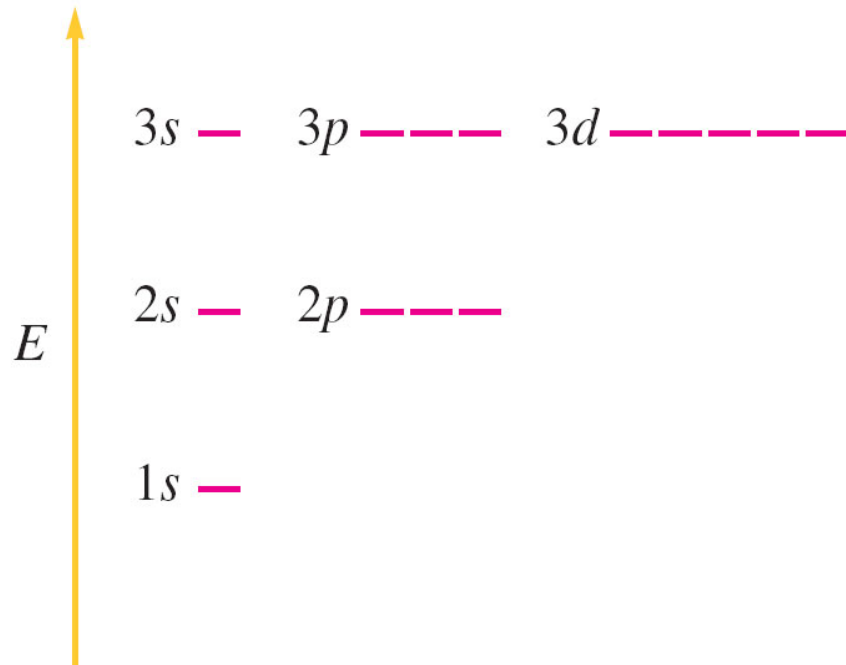


- Because most atoms contain many electrons occupying a number of different orbitals. If the orbitals are superimposed, then together they make up a 'spherical' shape.

Orbital Energies for the Hydrogen Atom

For the **H atom**:

- The orbital energy is determined by its n value.
- All orbitals with the same value of n have the same energy – they are **degenerate**.



- Hydrogen's single electron can occupy any of its orbitals.
- **Ground state** – the lowest energy state (1s orbital).
- **Excited state** – higher energy orbital which electron can occupy, but energy must be put into the atom to make the transfer.

Summary of the H Atom

- In the quantum (wave) mechanical model, the electron is viewed as a standing wave.
- A series of wave functions (orbitals) describe the possible energies and spatial distributions available to the electron.
- In agreement with the Heisenberg uncertainty principle, the model cannot specify the detailed electron motions.
- The square of the wave function represents the probability distribution of the electron in that orbital. Probability distribution gives electron density maps (a picture of the orbitals).

Summary of the H Atom

- The size of an orbital is defined as the surface that contains 90% of the total electron probability.
- H atoms have many types of orbitals. In the ground state, the single electron resides in the 1s orbital. The electron can be excited to higher-energy orbitals, if energy is put into the atom.

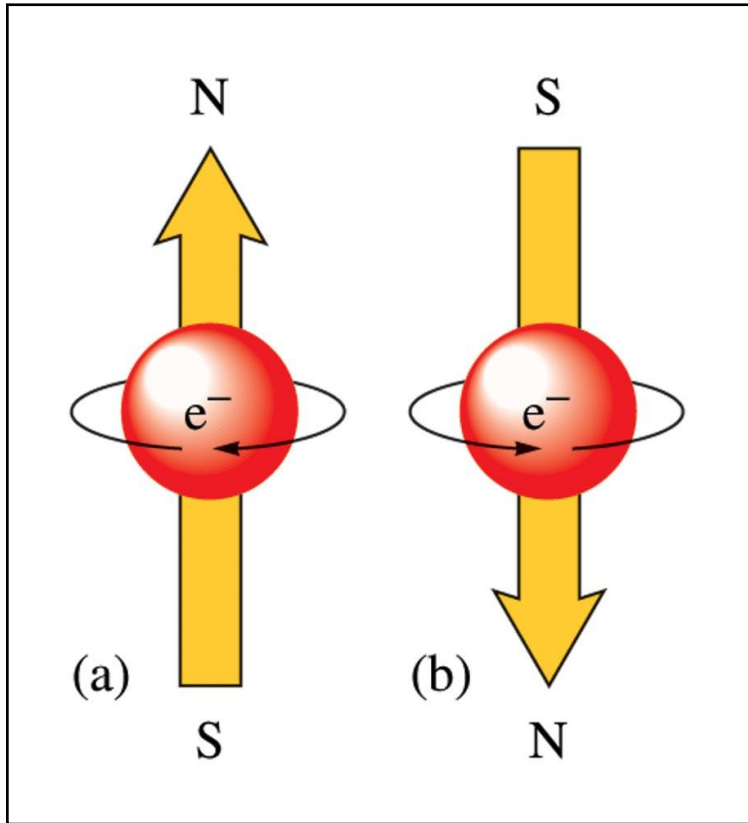
Polyelectronic atoms

Polyelectronic Atoms

Polyelectronic atoms are atoms with more than one electron.

Electron Spin and Pauli Principle

4th Quantum number - Electron Spin



Re-examination of emission line for polyelectronic elements often revealed two very closely spaced lines, where originally only one was observed. In 1925, this led Dutch physicists (Uhlenbeck and Goudsmit) to conclude that electrons have an intrinsic property – spin.

$$m_s = +\frac{1}{2} \text{ or } -\frac{1}{2}$$

A fourth quantum number exists and is the spin magnetic quantum number (m_s)

Electrons in Orbitals

- **Pauli Exclusion Principle**
 - No two electrons can have the same set of four quantum numbers (n , l , m_l and m_s)
 - Orbitals can hold only two electrons and they must have opposite spins.
 - Electron spin quantum number (m_s) – can be $+\frac{1}{2}$ or $-\frac{1}{2}$.



Wolfgang Pauli
Theoretical
physicist

Exercise

- Are the following sets of quantum numbers valid? Why or why not?

a) $n = 2, l = 1, m_l = -2, m_s = -1/2$

b) $n = 3, l = 2, m_l = -2, m_s = +1/2$

c) $n = 2, l = 1, m_l = 0, m_s = -1/2$

Exercise

- How many electrons can be described by the following quantum numbers?

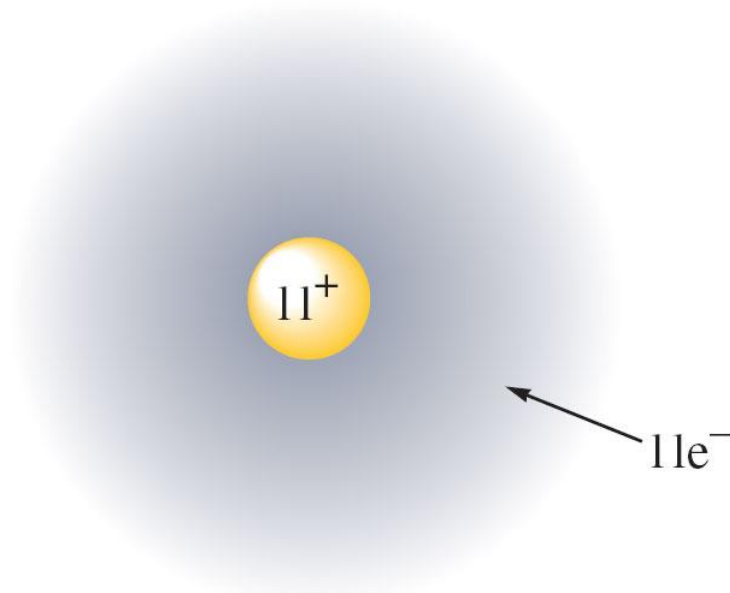
$$n = 3, m_l = -1, m_s = -1/2$$

Polyelectronic Atoms

- Polyelectronic atoms are atoms with more than one electron.
- Electron correlation problem:
 - Since the electron pathways are unknown, the electron repulsions cannot be calculated exactly.
- When electrons are placed in a particular quantum level, they “prefer” the orbitals in the order s , p , d , and then f .

Mnemonic : Some Poor Dumb Fool

Shielding and Penetration

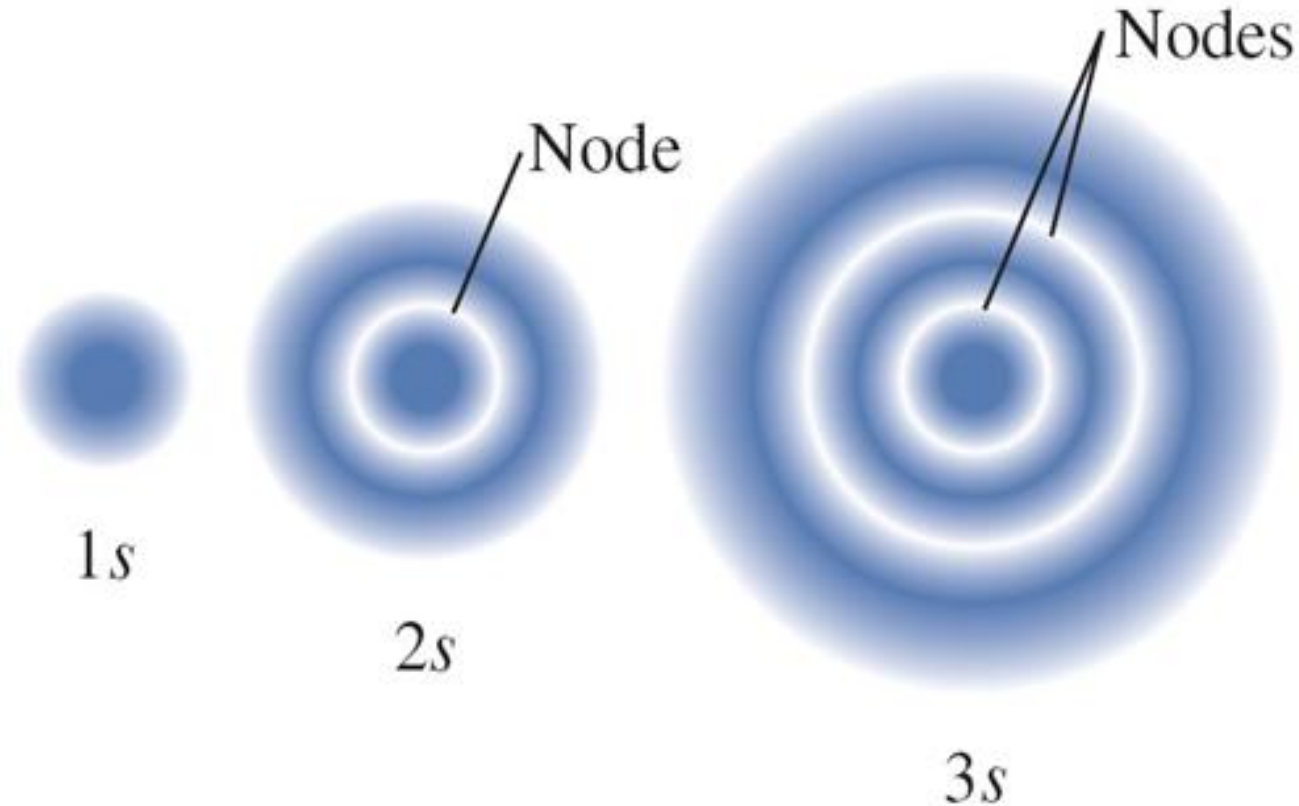


- In the sodium atom, the outermost electron:
 - does not feel the full nuclear charge,
 - it is **shielded** by the other electrons,
 - is not bound as tightly to the nucleus as it would be if the other electrons were not present.

Shielding and Penetration

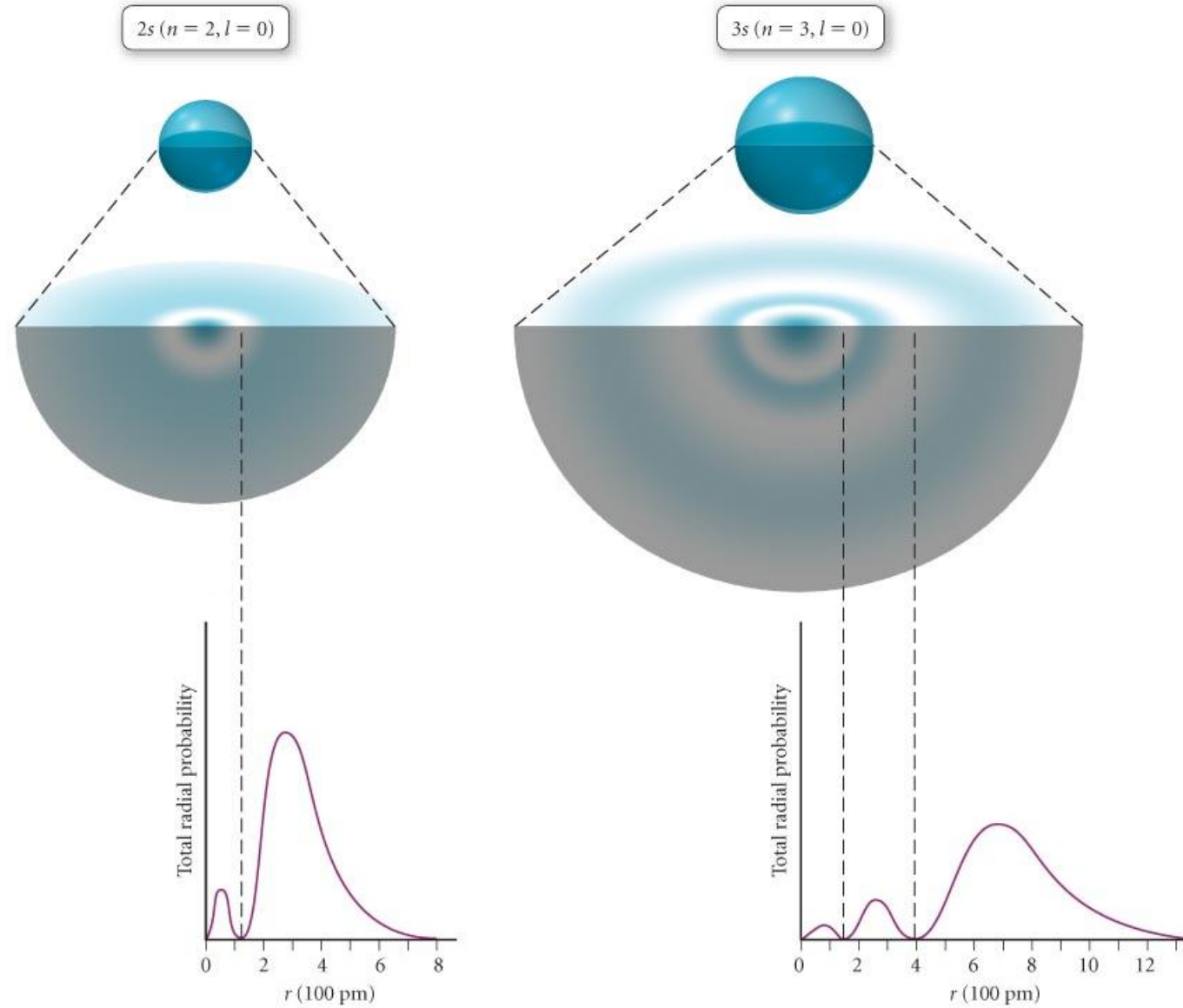
- Electrons in orbitals closer to the nucleus **screen or shield** the nucleus from electrons farther away.
- In effect, the screening electrons reduce the effectiveness of the nucleus in attracting the particular more distant electron; they reduce the **effective nuclear charge**.

The Radial Probability Distribution of the 1s, 2s and 3s Orbital

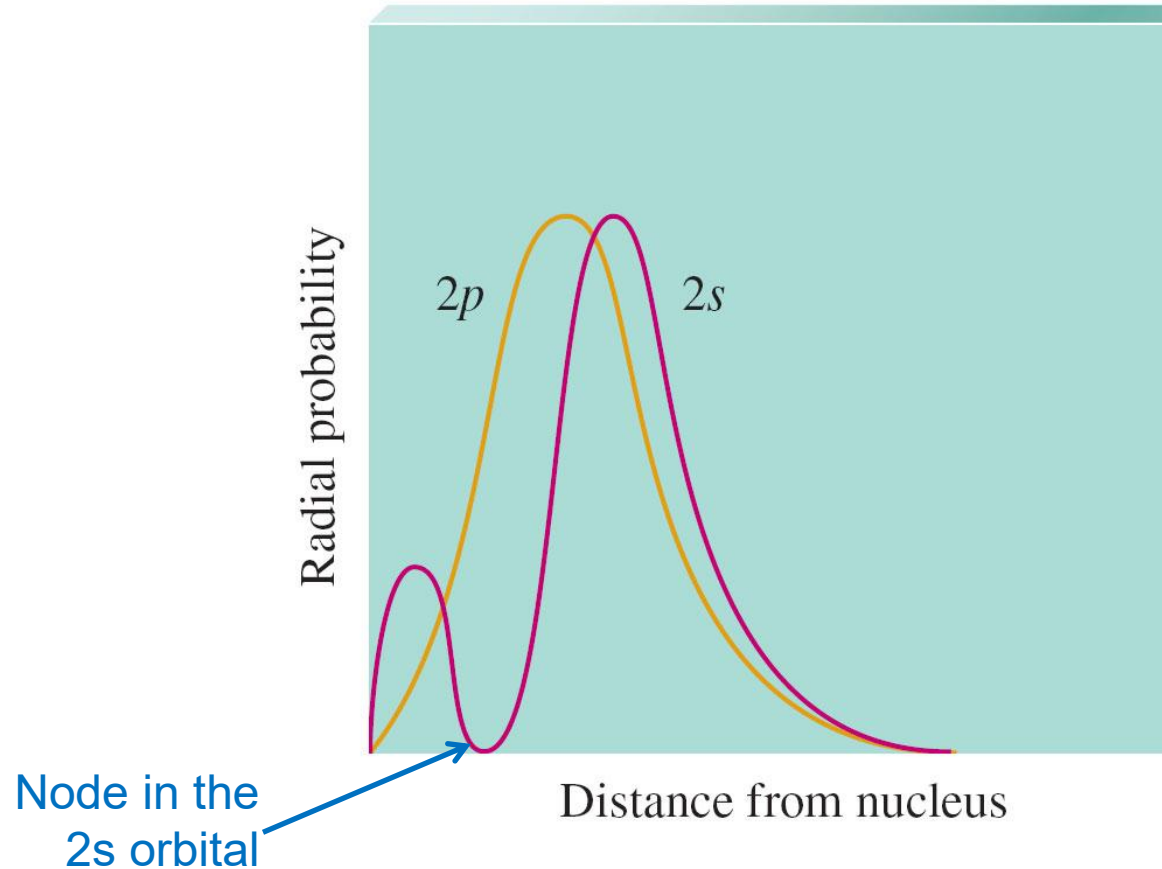


- Although electrons spend most of their time in the space further away from the nucleus, they spend a small but **SIGNIFICANT** amount of time very near the nucleus.
- They don't spend any time in the nodes.

The 2s and 3s Orbitals

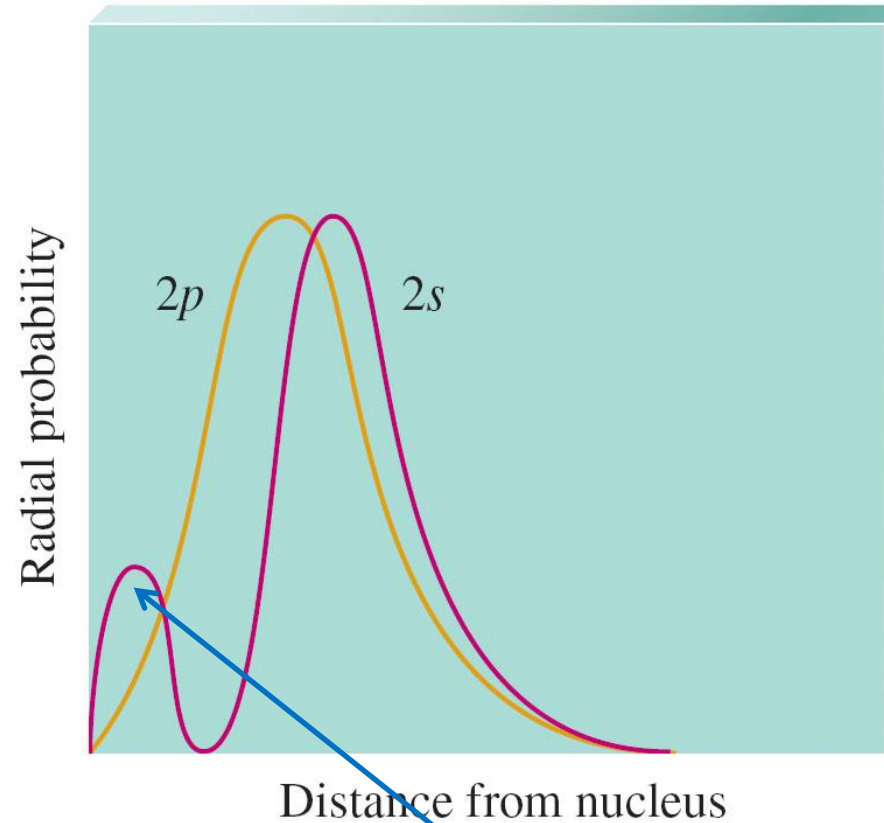


Shielding and Penetration



The electron in the 2s orbital spends a small but **significant time** very near the nucleus

Shielding and Penetration

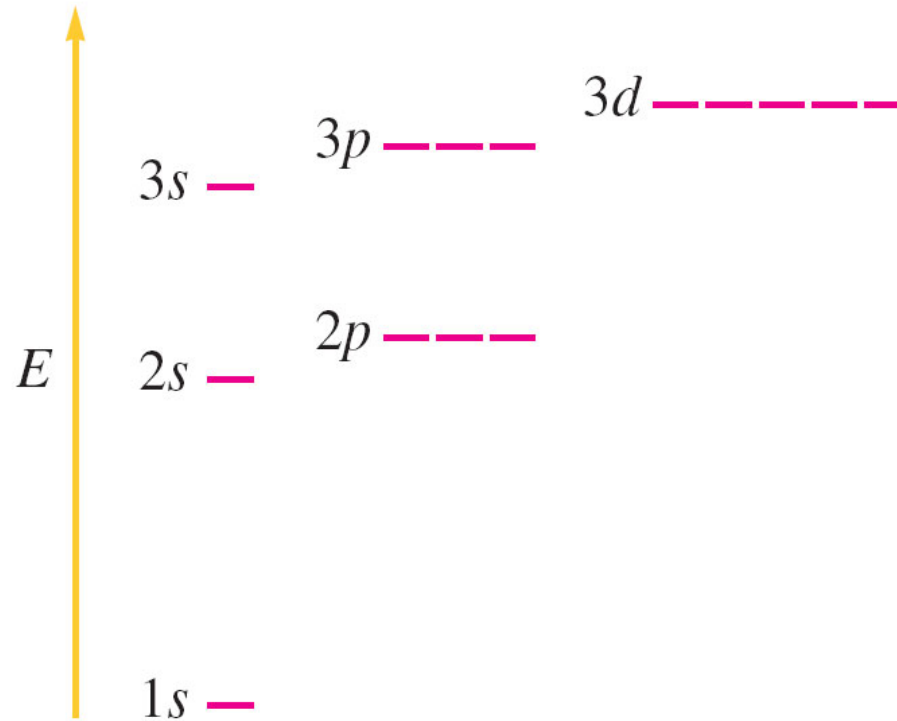


The electron in the 2s orbital **penetrates** to the nucleus more than the 2p orbital so the 2s is **lower** in energy!

Penetration Effect - Summary

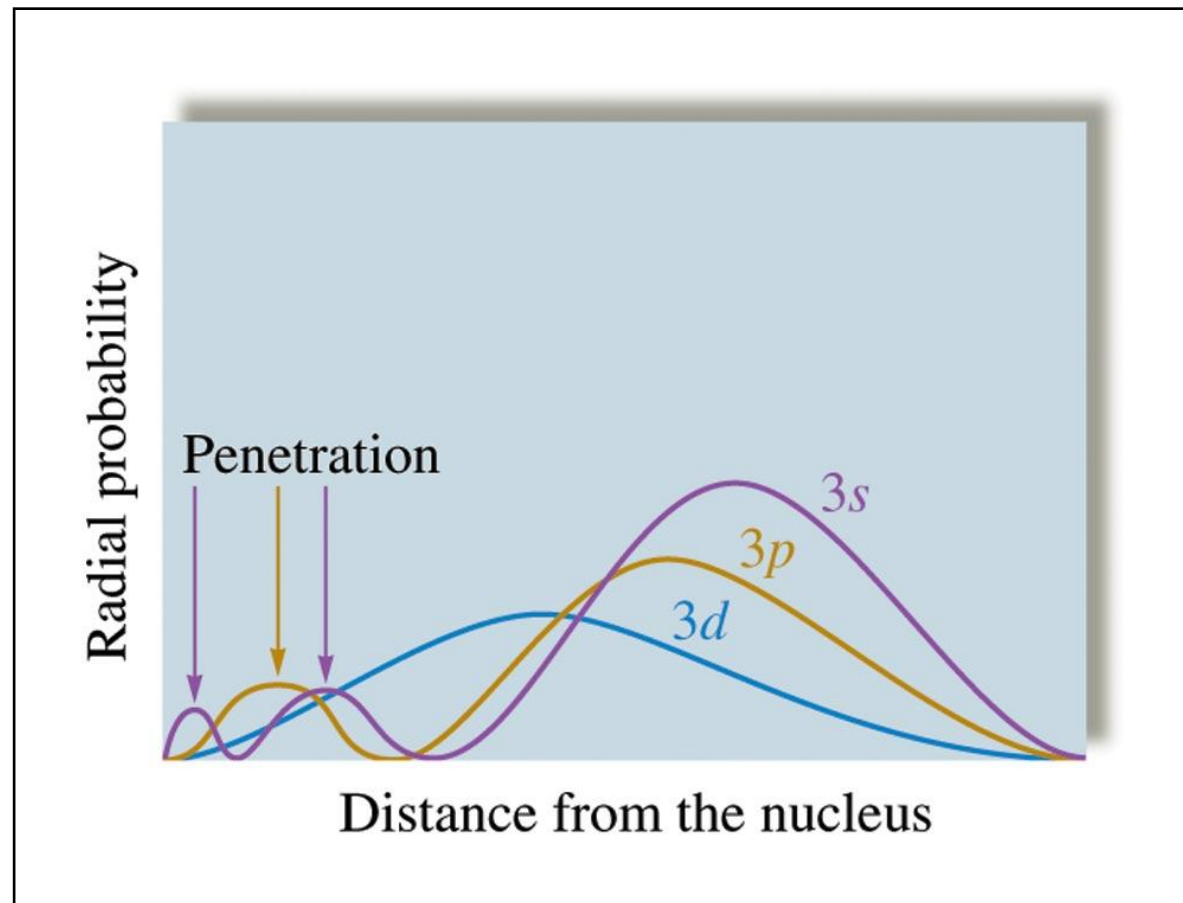
- A $2s$ electron penetrates to the nucleus more than an electron in the $2p$ orbital.
- This causes an electron in a $2s$ orbital to be attracted to the nucleus more strongly than an electron in a $2p$ orbital.
- Thus, the $2s$ orbital is **lower** in energy than the $2p$ orbitals in a polyelectronic atom.

Polyelectronic Atoms and Orbital Energy Levels



Because of shielding and penetration, the orbitals of polyelectronic atoms are **not** degenerate (for a given n)

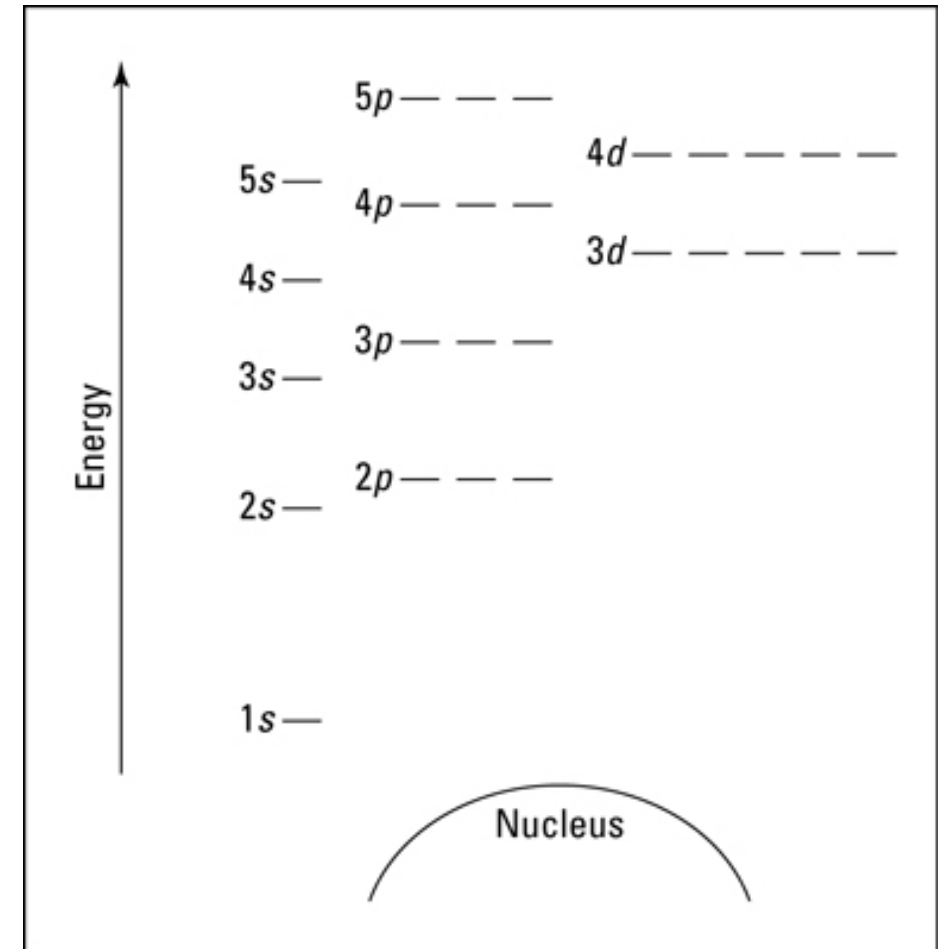
A Comparison of the Radial Probability Distributions of the $3s$, $3p$, and $3d$ Orbitals



Higher Energy Levels

Notice the following:

1. Because of penetration, subshells within an energy level are not degenerate.
2. Penetration of the fourth and higher energy levels is so strong that their *s* subshells are lower in energy than the *d* subshells of the next lower energy level, (example: 4*s* vs 3*d*).
3. The energy difference between levels becomes smaller for higher energy levels (and can cause anomalous electron configurations for certain elements).



Quantum Mechanical Model of the Atom

Quantum Number	Symbol	Values	Function
Principal quantum number (<i>shell</i>)			
	l		
	m_l		
Electron spin quantum number			

Angular Momentum and Magnetic Quantum Numbers

l value:	0	1	2	3
l letter:				
m_l value(s):				
maximum # electrons				

Quantum Mechanical Model of the Atom

Figure 5.15 *Figure It Out* :

What are the differences between 1s, 2s and 3s orbitals?

Quantum Mechanical Model of the Atom

Question: If $l = 1$ (or p), what are the possibilities for m_l ?

Quantum Mechanical Model of the Atom

Big Idea Question 5.5

What are the possible values of n , l , m_l for an electron in a 4d orbital?

Quantum Mechanical Model of the Atom

Activity: Answer the following 6 questions.

1) What is the maximum number of electrons that an orbital can hold?

2) Which orbital is spherically symmetrical?

3) How many orientations in space can p orbitals have?

a. 1

b. 3

c. 5

d. 7

4) All the orbitals in a given shell have the same value for which quantum number?

· Principle, n

· Angular, l

· Magnetic, m_l

· Spin, m_s

Quantum Mechanical Model of the Atom

5) For each set, state which quantum number is invalid and why.

$n = 1, l = 1, m_l = 0, m_s = +1/2$	
$n = 2, l = 1, m_l = -1, m_s = +3/2$	
$n = 3, l = 1, m_l = 2, m_s = -1/2$	
$n = 0, l = 1, m_l = 0, m_s = +1/2$	

6) True or False: The **Pauli Exclusion Principle** states that two electrons can have precisely the same four quantum numbers. Support your answer.

Aufbau Principle and the Periodic Table

Aufbau Principle and the Periodic Table

Aufbau (German for “building up”)

H (Z = 1)

He (Z = 2)

Li (Z = 3)

Be (Z = 4)

B (Z = 5)

As we go across the periodic table, the atomic number increases which is the same as saying the number of protons increases.

Aufbau Principle and the Periodic Table

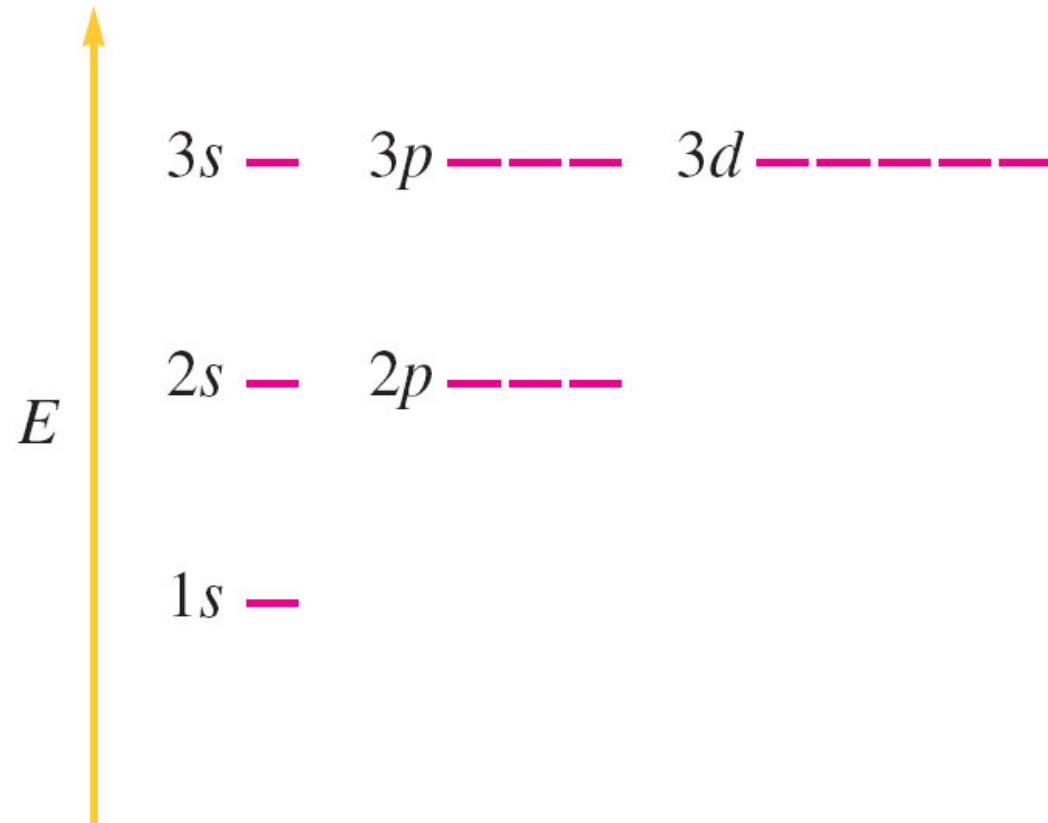
Aufbau (German for “building up”)



$$Z = 1$$

As the number of protons increases so does the number of electrons. We need to figure out where they go.

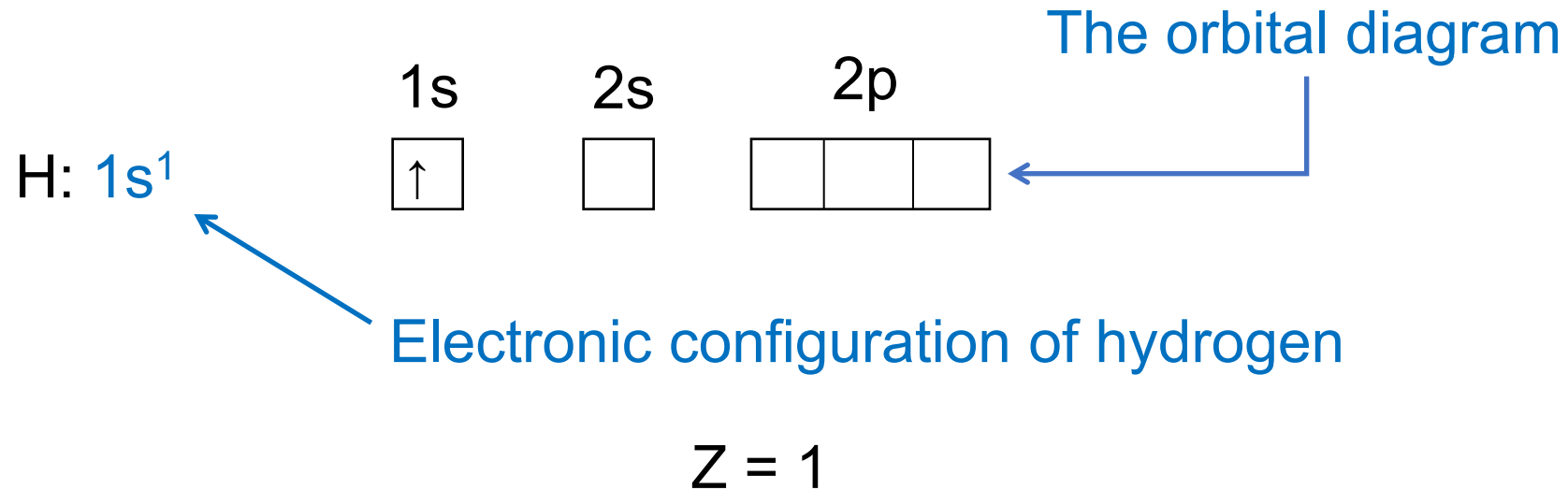
Energy Levels of Orbitals



The lower the principal quantum number, the lower the energy, thus the $1s$ orbital is the lowest in energy.

Aufbau Principle and the Periodic Table

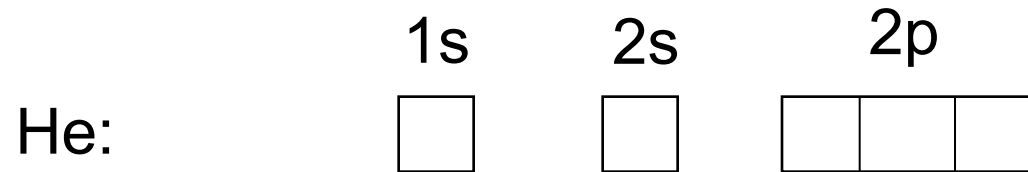
Aufbau (German for “building up”)



We add electrons in the lowest energy orbitals first!

Aufbau Principle and the Periodic Table

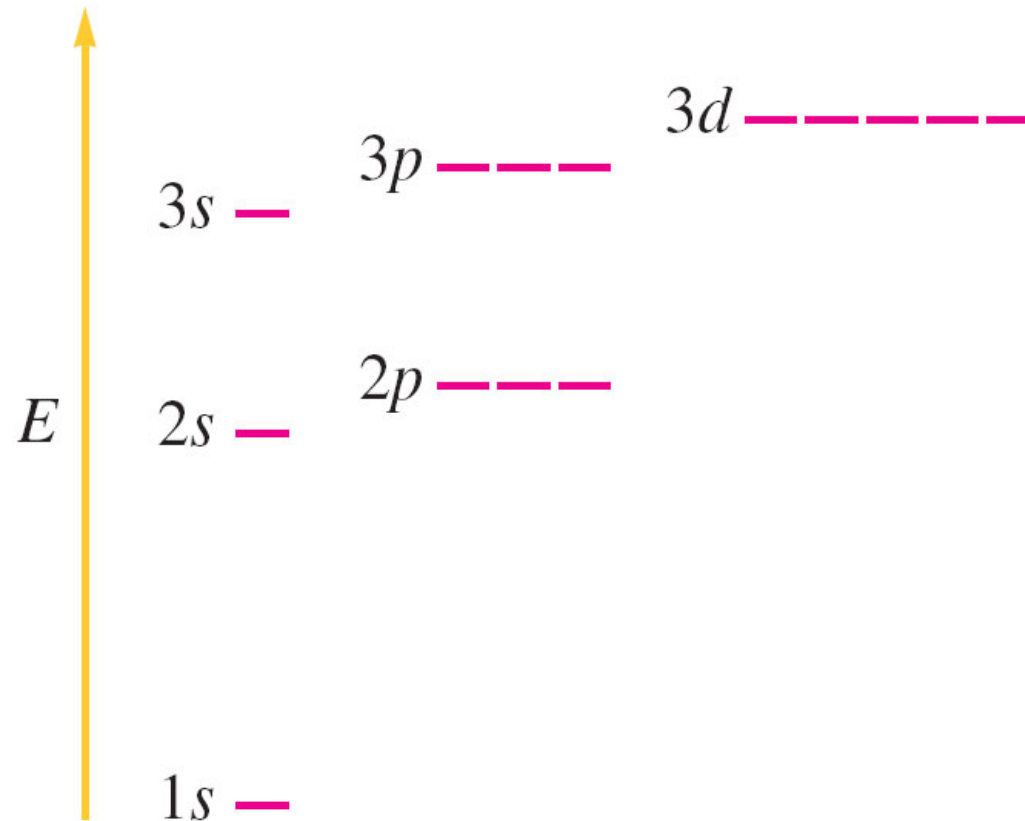
Aufbau (German for “building up”)



$$Z = 2$$

The next element to consider is helium.

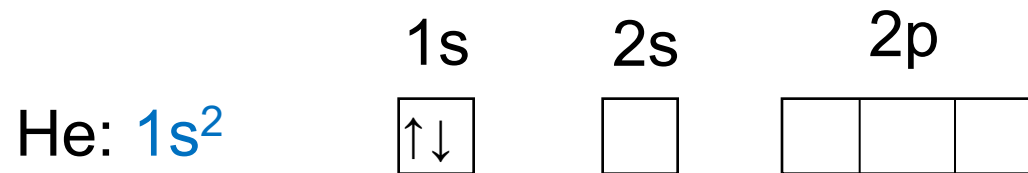
Polyelectronic Atoms and Orbital Energy Levels



Because of shielding and penetration, the orbitals of polyelectronic atoms are not degenerate (for a given n).

Aufbau Principle and the Periodic Table

Aufbau (German for “building up”)



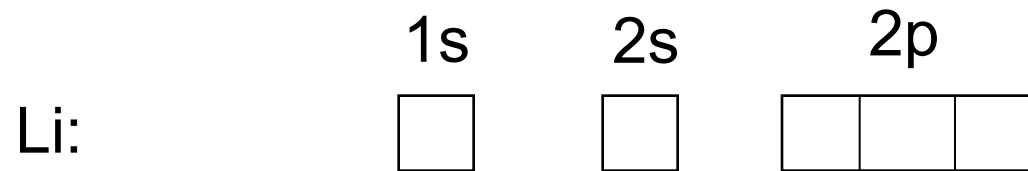
$$Z = 2$$

Electrons in the same orbital must have opposite spins: **Pauli Exclusion Principle**.

Configuration of He is a noble gas configuration.

Aufbau Principle and the Periodic Table

Aufbau (German for “building up”)

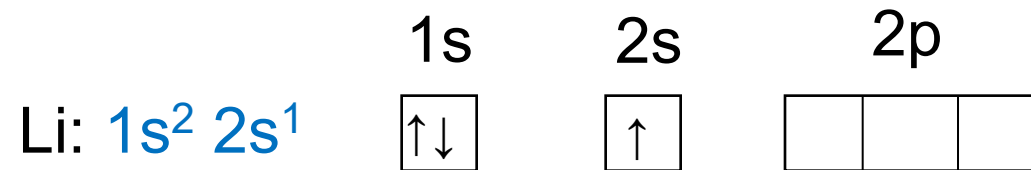


$$Z = 3$$

We continue to fill orbitals as we progress to elements with higher values of Z .

Aufbau Principle and the Periodic Table

Aufbau (German for “building up”)

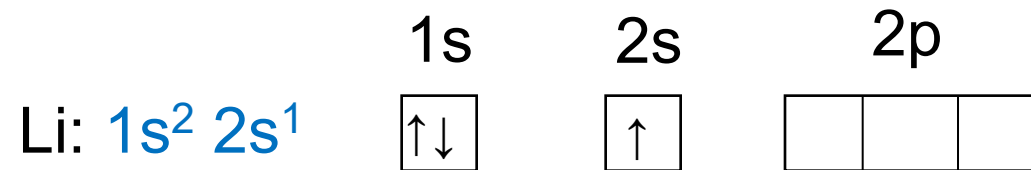


$$Z = 3$$

We continue to fill orbitals as we progress to elements with higher values of Z .

Aufbau Principle and the Periodic Table

Aufbau (German for “building up”)



Configuration of Li can also be:

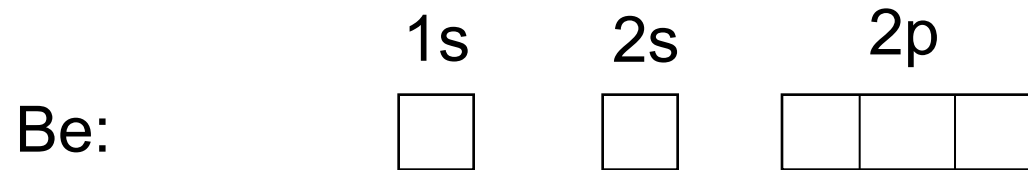


Core electrons
(innermost electrons)

Valence electrons
(outermost electrons)

Aufbau Principle and the Periodic Table

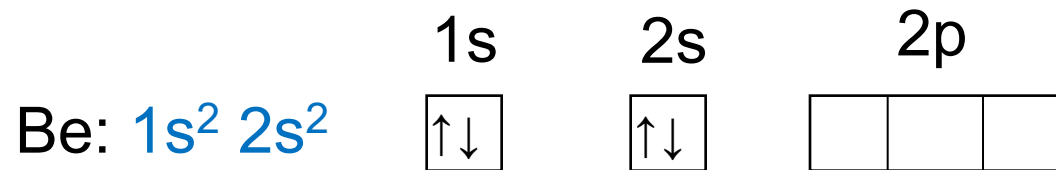
Aufbau (German for “building up”)



$$Z = 4$$

Aufbau Principle and the Periodic Table

Aufbau (German for “building up”)

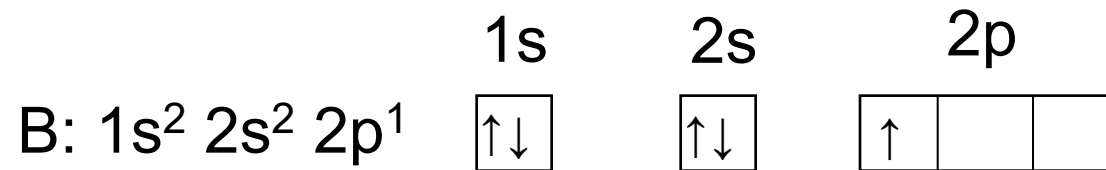


$$Z = 4$$

The energy required to pair the electrons is less than the energy between the 2s and 2p orbitals.

Aufbau Principle and the Periodic Table

Aufbau (German for “building up”)

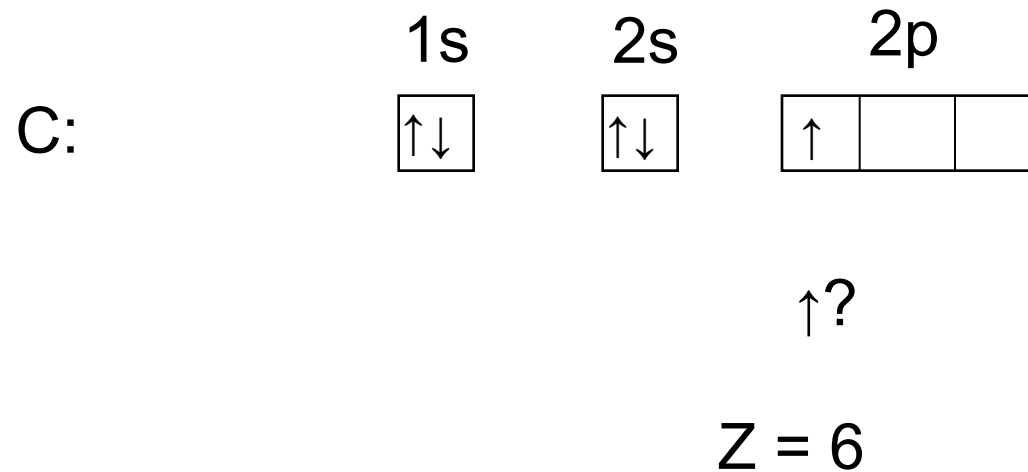


$$Z = 5$$

The next orbital to be filled is the next one higher in energy: the 2p.

Aufbau Principle and the Periodic Table

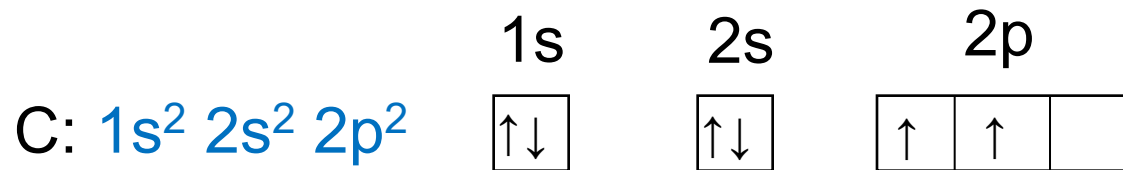
Aufbau (German for “building up”)



Where do we put the next electron?

Aufbau Principle and the Periodic Table

Aufbau (German for “building up”)

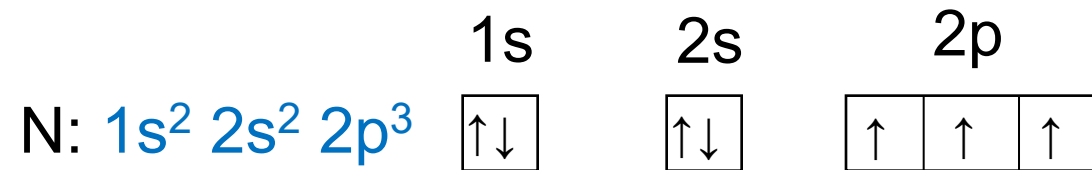


$$Z = 6$$

Hund's rule: Lowest energy configuration is the one having the maximum number of unpaired electrons in a set of degenerate orbitals.

Aufbau Principle and the Periodic Table

Aufbau (German for “building up”)

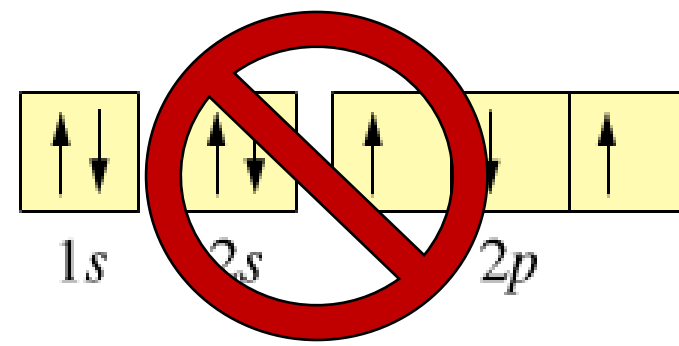
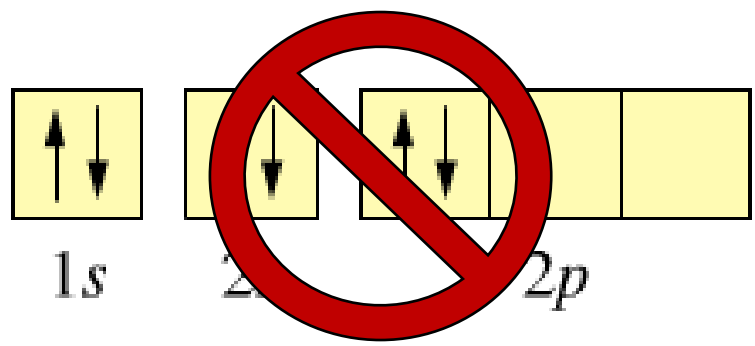
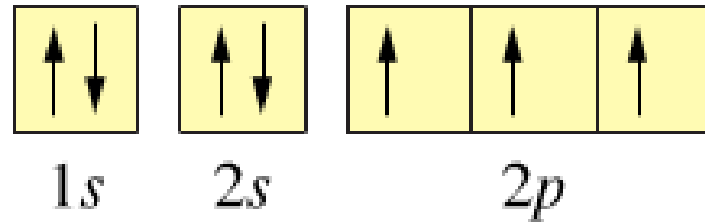


$$Z = 7$$

Hund's rule: Lowest energy configuration is the one having the maximum number of unpaired electrons in a set of degenerate orbitals.

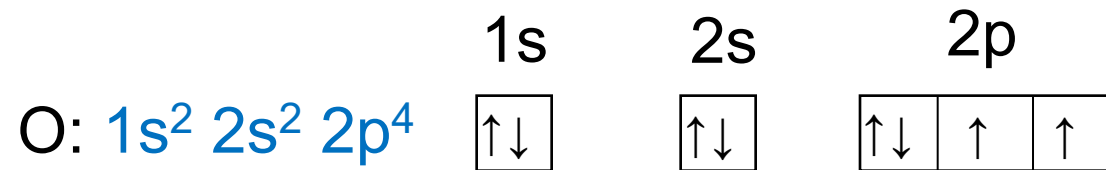
Electron Configuration

Hund's Rule: within a subshell, the e- first occupy the empty degenerate orbitals before pairing.



Aufbau Principle and the Periodic Table

Aufbau (German for “building up”)

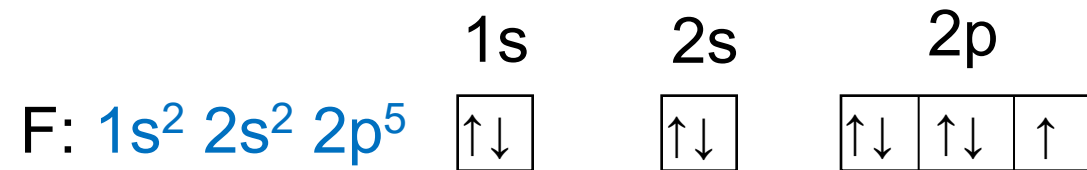


$$Z = 8$$

Once all the empty orbitals are filled with unpaired electrons, they can then be paired.

Aufbau Principle and the Periodic Table

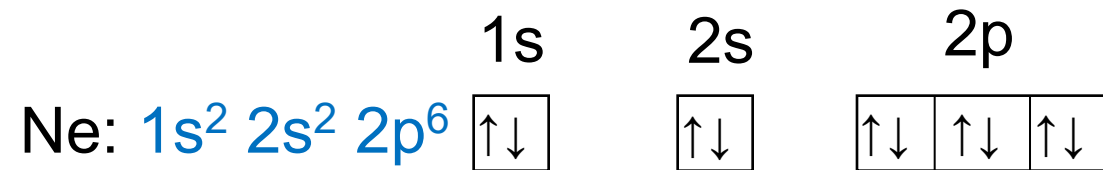
Aufbau (German for “building up”)



$$Z = 9$$

Aufbau Principle and the Periodic Table

Aufbau (German for “building up”)

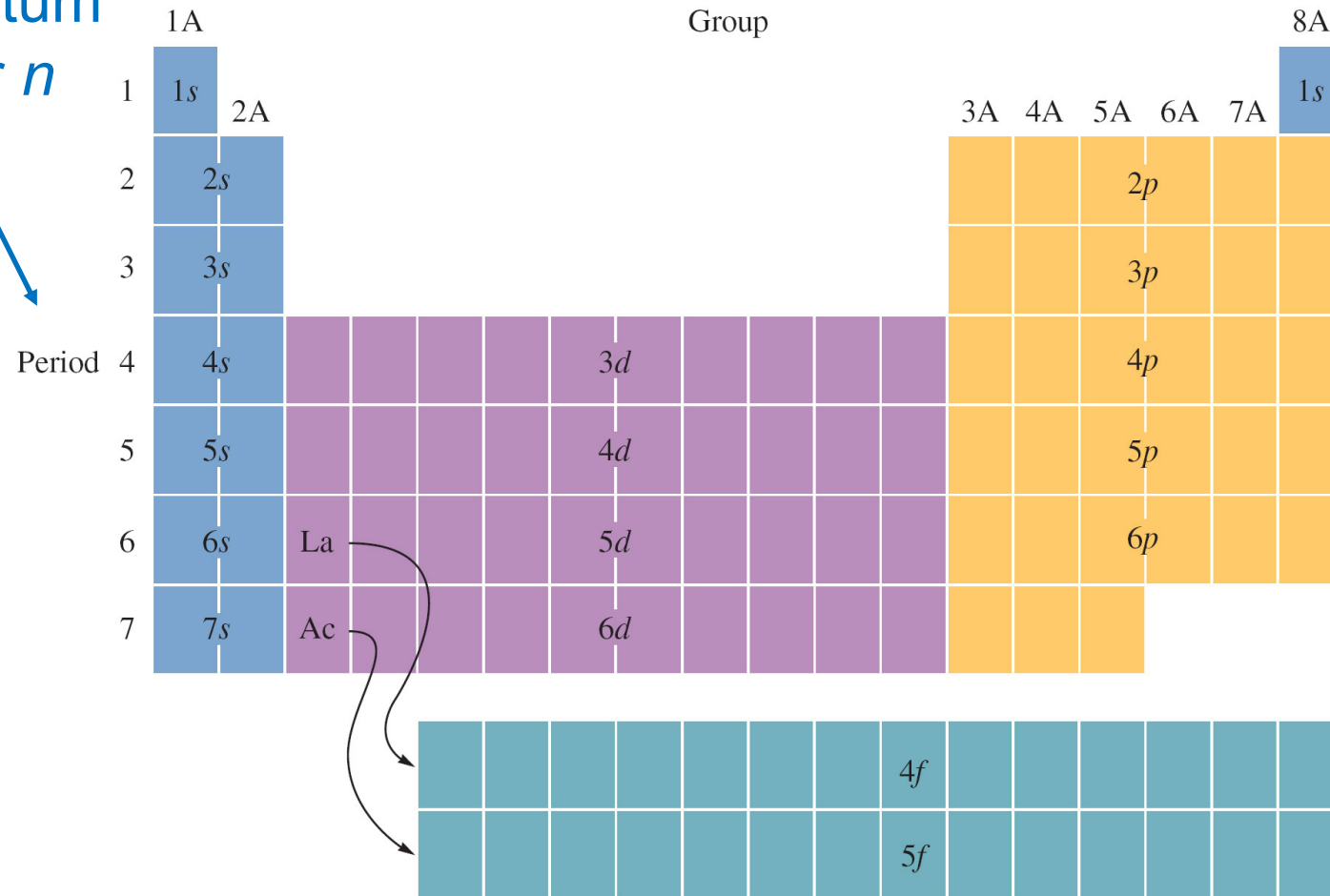


$$Z = 10$$

Neon has a noble gas configuration.

Electronic Configuration and the Periodic Table

Corresponds to quantum number n

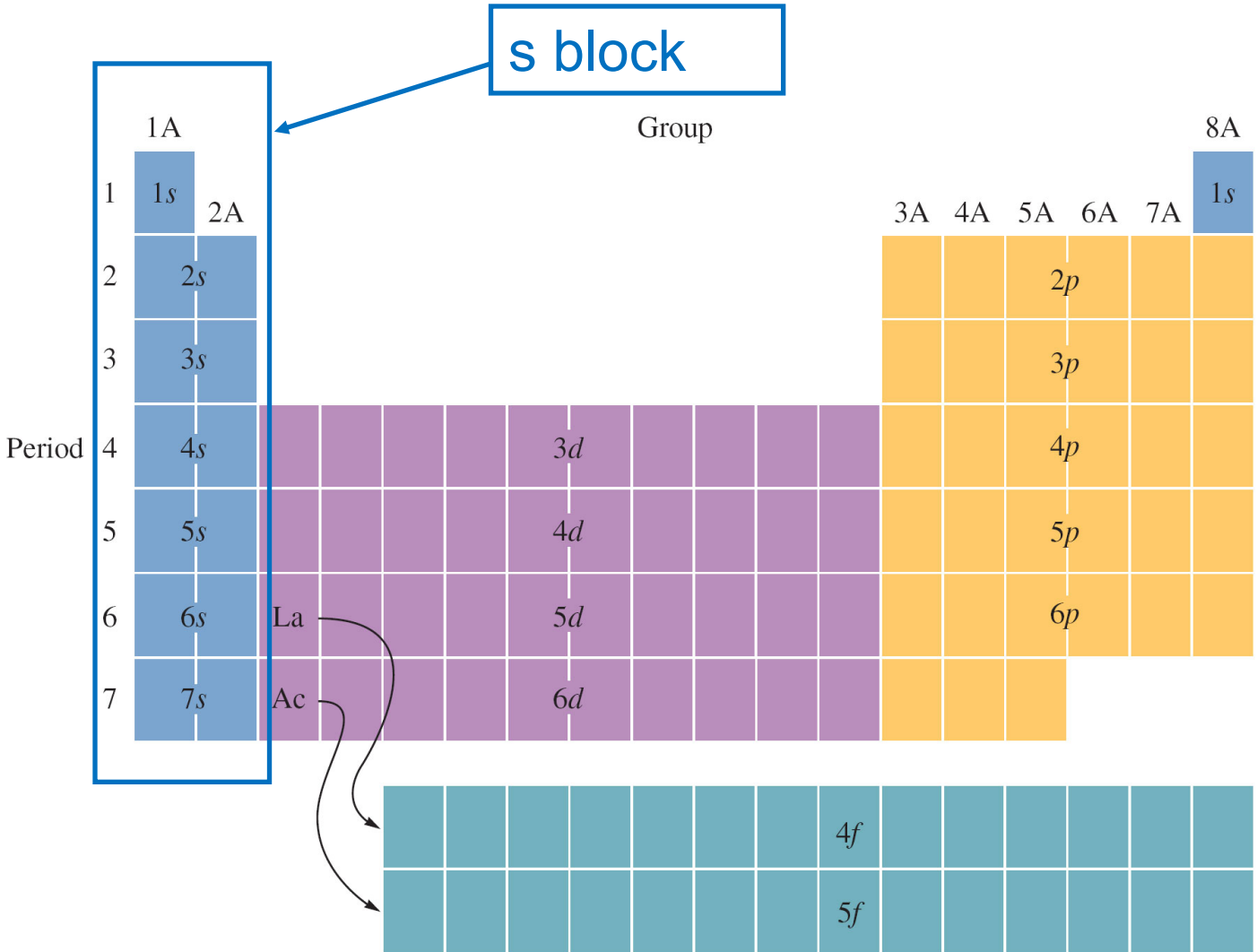


Electronic Configuration and the Periodic Table

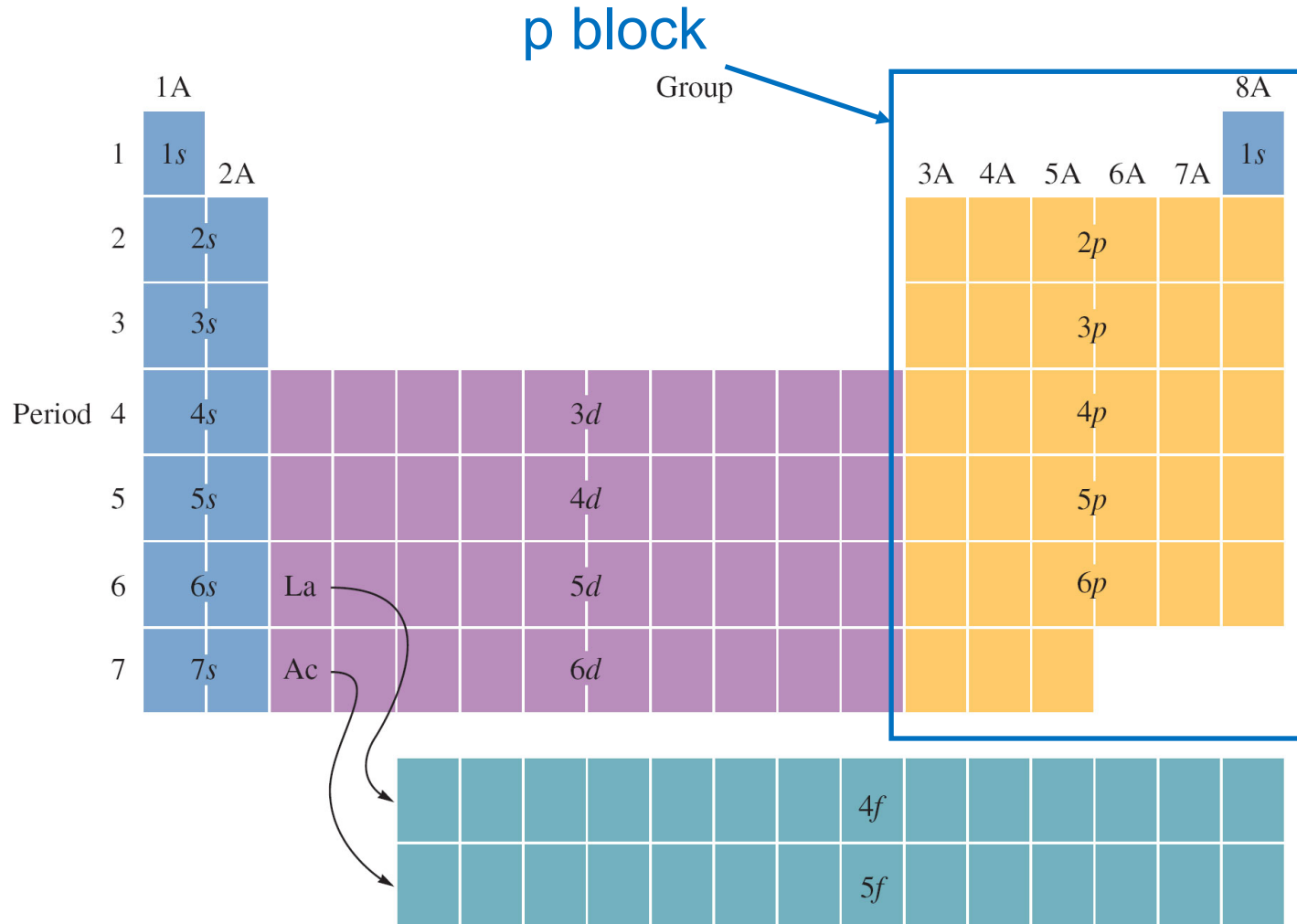
Group

	1A											8A				
1	1s	2A											1s			
2	2s											3A 4A 5A 6A 7A				
3	3s											2p 3p				
4	4s									3d			4p			
5	5s									4d			5p			
6	6s	La									5d			6p		
7	7s	Ac									6d					
												4f				
												5f				

Electronic Configuration and the Periodic Table

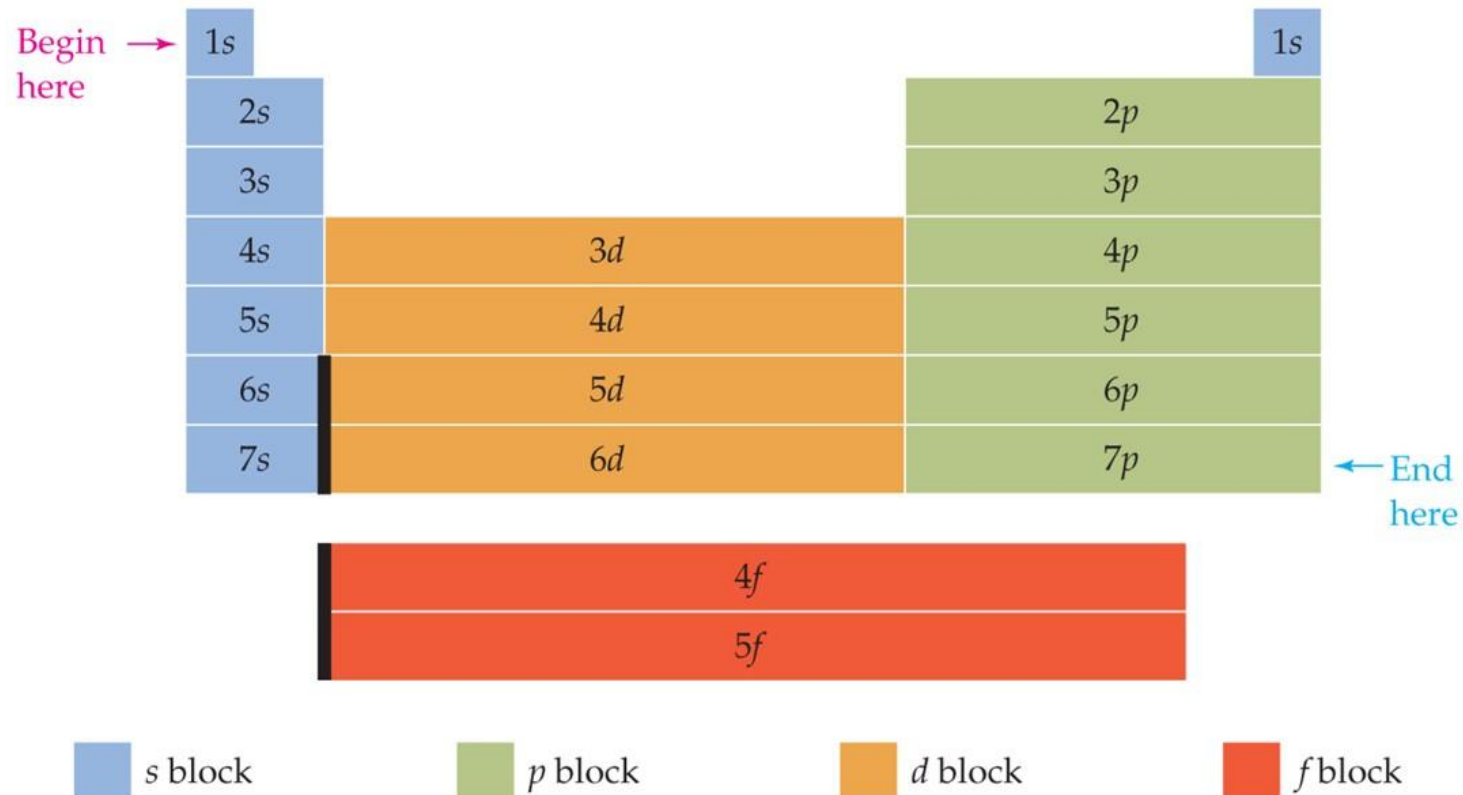


Electronic Configuration and the Periodic Table

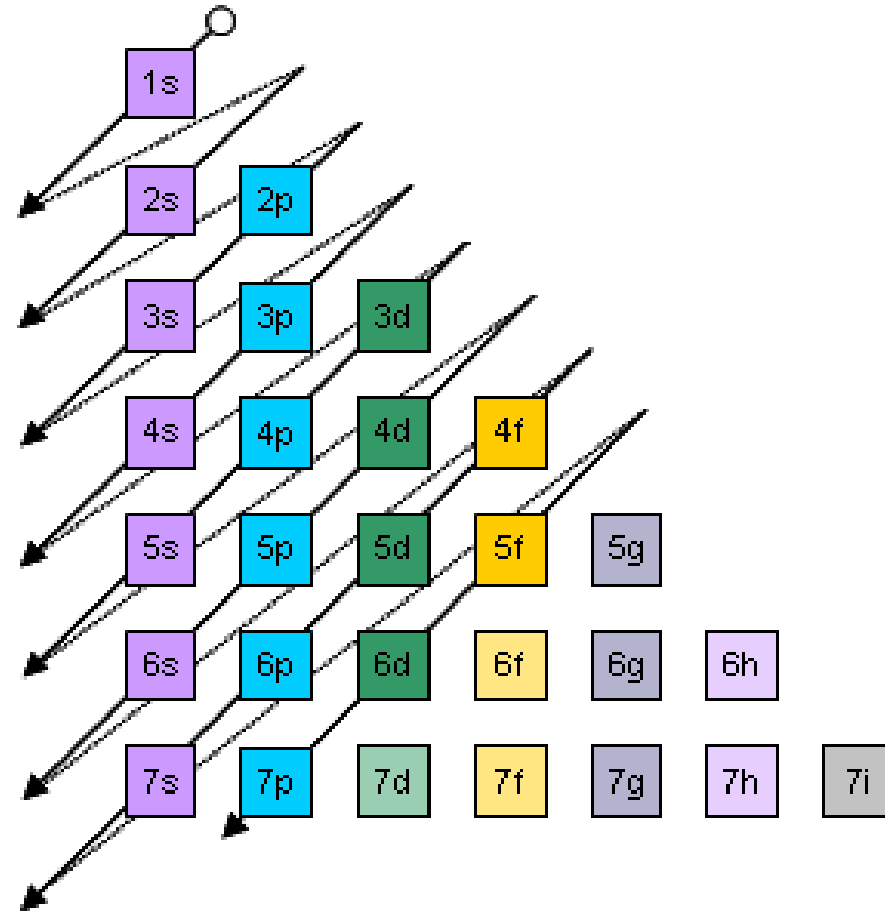


Electron Configurations and the Periodic Table

The arrangement of the periodic table provides a method for remembering the order of orbital filling. Beginning at the top left and moving across successive rows, the order is $1s \rightarrow 2s \rightarrow 2p \rightarrow 3s \rightarrow 3p \rightarrow 4s \rightarrow 3d \rightarrow 4p$ and so on.



Mnemonic - Order of Orbital Fill



Order: 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s, 5f, 6d, etc.

Electronic Configuration and the Periodic Table

What is the electronic configuration of calcium?

The periodic table is color-coded and labeled as follows:

- Alkaline earth metals:** Group 2 (2A), highlighted in purple.
- Transition metals:** Groups 3-10, highlighted in yellow.
- Halogens:** Group 17 (7A), highlighted in cyan.
- Noble gases:** Group 18 (8A), highlighted in blue.

1 H	2 He											13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg	3	4	5	6	7	8	9	10	11	12	13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
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
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
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
87 Fr	88 Ra	89 Ac†	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Uub	113 Uut	114 Uuq	115 Uup			118 Uuo

Electronic Configuration and the Periodic Table

What is the electronic configuration of titanium?

The periodic table is color-coded and labeled with groups and periods. The groups are labeled as follows: 1A (Alkali metals), 2A (Alkaline earth metals), 3A-8A (Main groups), and 8A (Noble gases). The transition metals are labeled as such. The element titanium (Ti) is highlighted with a red circle.

1A	2A	Transition metals										3A	4A	5A	6A	7A	8A
1 H	2 He																
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg	3	4	5	6	7	8	9	10	11	12	13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
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
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
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
87 Fr	88 Ra	89 Ac†	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Uub	113 Uut	114 Uuq	115 Uup			118 Uuo

Electron Configuration

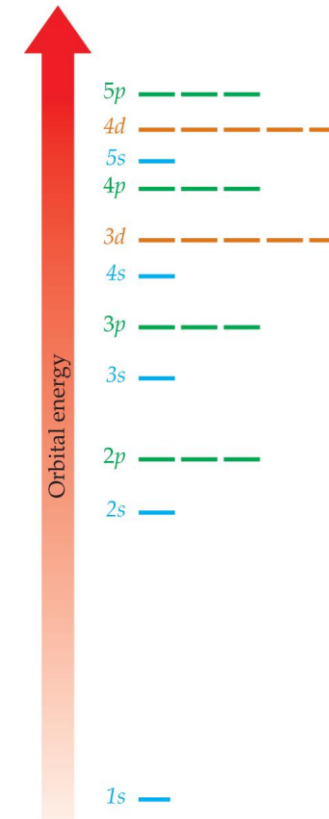
Rules of the Aufbau principle:

1. Fill lower-energy orbitals before higher-energy orbitals.
2. An orbital can hold only two electrons, which must have opposite spins (Pauli exclusion principle).
3. If two or more degenerate orbitals are available, follow Hund's rule.

Hund's Rule:

If two or more orbitals with the same energy (degenerate orbitals) are available, one electron goes into each until all are _____.

The electrons in the half-filled orbitals all have the _____ value of their spin quantum number (i.e. all have the same spin).



VIDEO: [Aufbau's Principle, Hund's Rule & Pauli's Exclusion Principle](#)

Exercise

Determine the expected electron configurations for each of the following.

a) S

b) Ba

c) Eu

Exercise

A ground-state electron in the hydrogen atom is given just enough energy to get to $n = 2$. Which orbital will the electron occupy?

- a) The $2s$ orbital.
- b) The $2p_x$ orbital.
- c) The $2p_y$ orbital.
- d) The $2p_z$ orbital.
- e) Each of the above orbitals is equally likely.

Exercise

Of the following elements, which has occupied d orbitals in its ground-state neutral atoms?

- a) Ba
- b) Ca
- c) Si
- d) P
- e) Cl