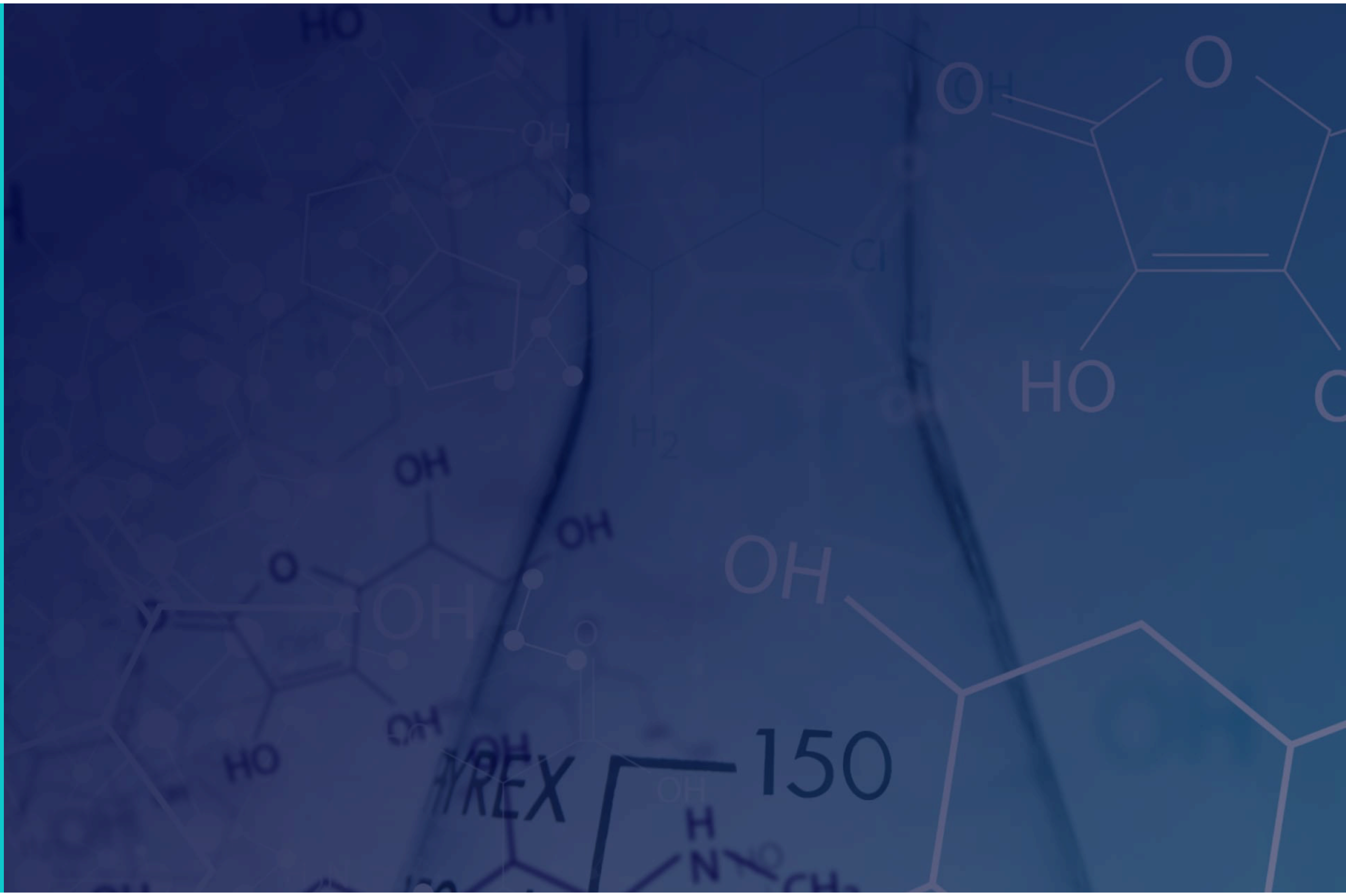

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CHM151Y

Summer 2025, Chapter 4 Notes

Table of Contents

4. Quantum Numbers and Electron Configurations

4.1 The Quantum Model & Orbitals

4.1.1

Quantum Numbers

Quantum numbers describe where electrons are positioned around atoms:

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

Rules:

n

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

- As n increases, energy and size of shell ________

WIZE TIP

n is also called the "principle quantum number"

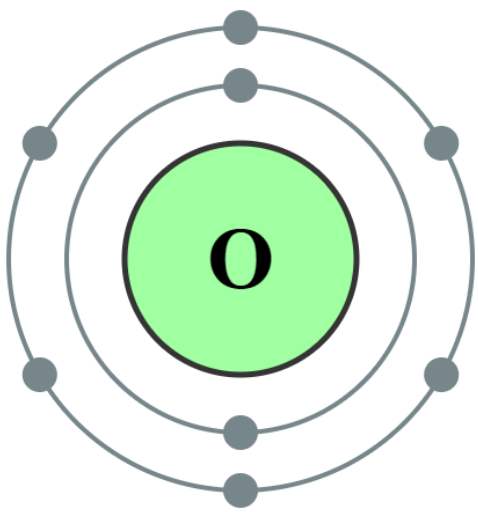


Photo by Greg Robson / CC BY

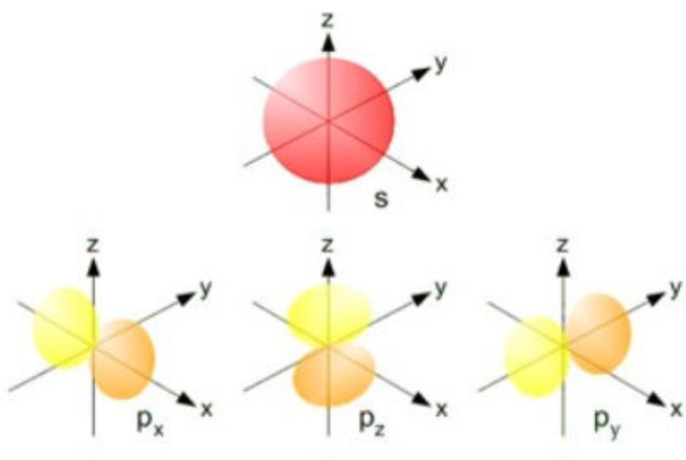
l

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

- l (orbital shapes) are described as:

subshell	s	p	d	f
l	0	1	2	3

Increasing energy →



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

WIZE CONCEPT

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

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

m_l

- Can be any integer from **–l to +l**
  Example: m$_1$ = (-l, ..., 0, ..., l)

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

m_s

- **Spin** of an electron

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

4.1.2

Quantum Numbers Summary

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

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

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, -$^1/_2$)

4.1.5

Example: Defining Orbitals from Quantum Numbers

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

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

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

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

4.1.6

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, $-{}^1/_2$)

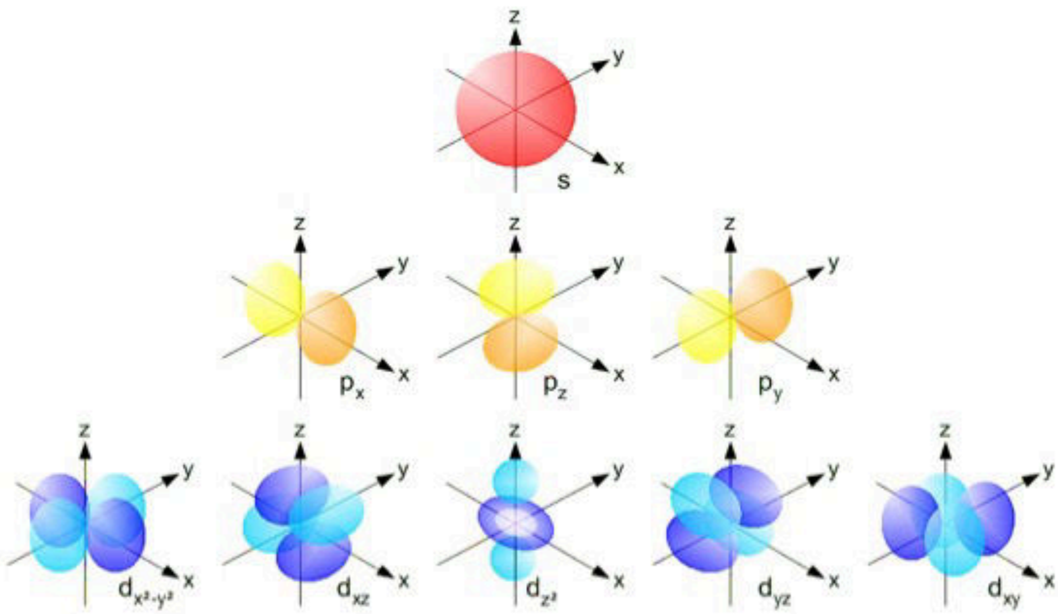
e) (3, 2, 3, $-{}^1/_2$)

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

4.2 Shapes of Atomic Orbitals

4.2.1

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)

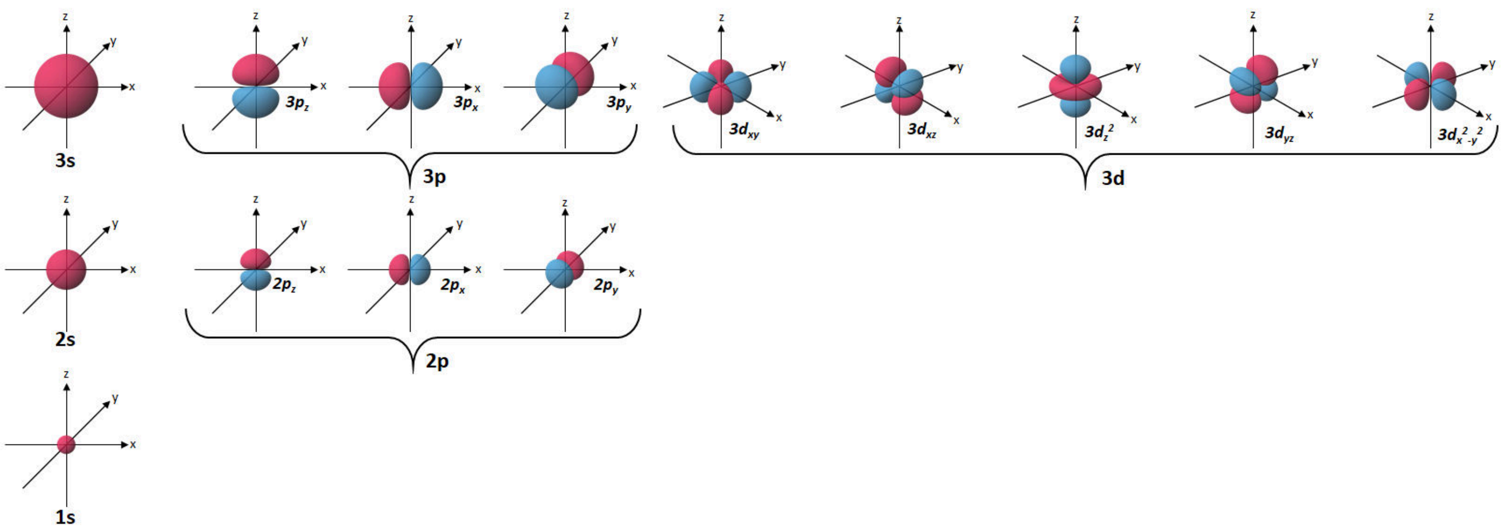
- There are **five d orbitals** in a subshell. The shapes of the d orbitals are more complicated than their s and p counterparts.
    - Three orbitals look like 3D cloverleaves, each lying in a plane with the lobes pointed between the axes (d_{xy}, d_{xz} and d_{yz}).
    - A fourth orbital is also a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

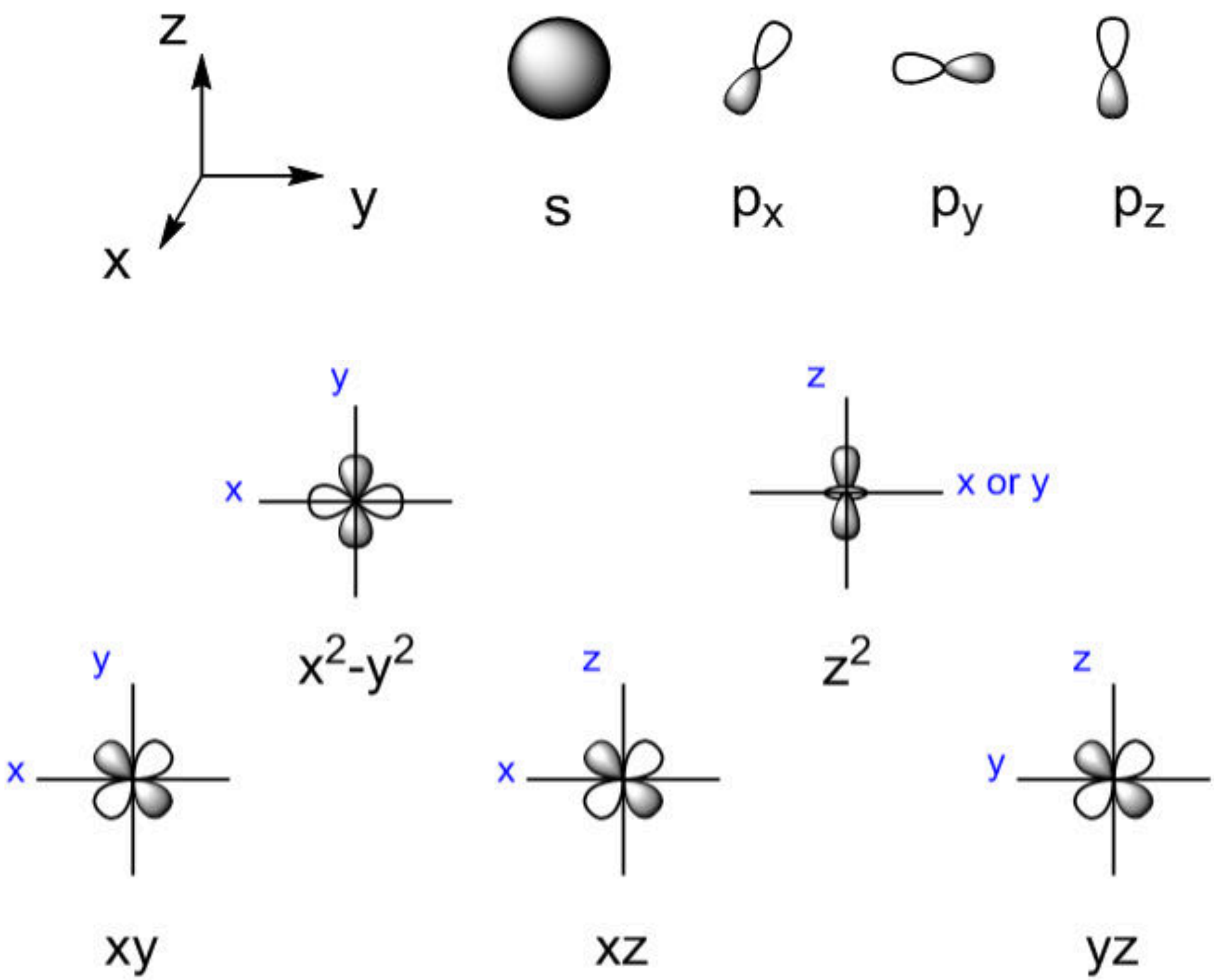
- As the **number of orbitals increases**, so does their **complexity**
  Example: d orbitals are more complex than p orbitals



4.2.2

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!

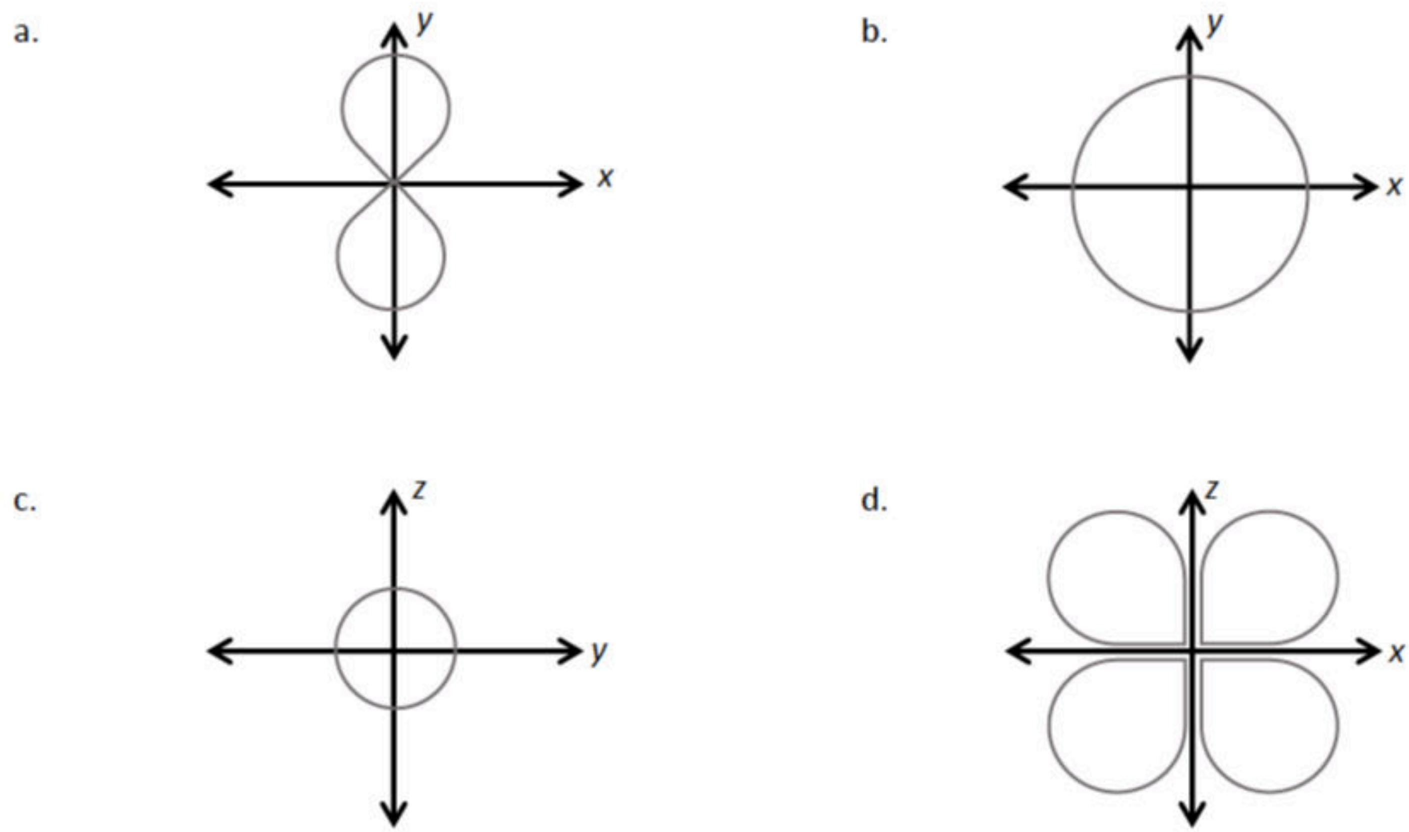
Also, notice how every other lobe is colored in gray. This gray and white coloring is indicating a "positive" lobe and a "negative" lobe. If you were asked to draw one of these orbitals on an exam, you would be expected to color in the lobes as well (or indicate which are + and which are -).

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

4.2.3

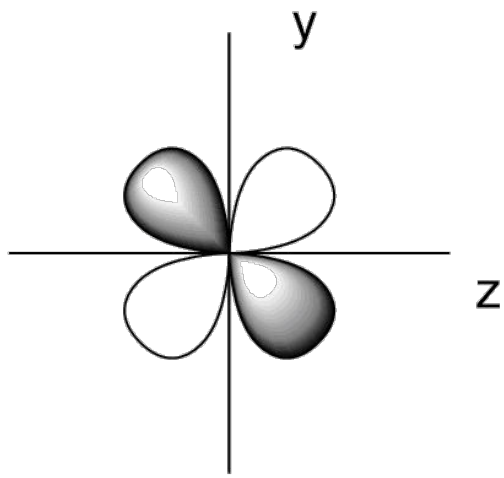
Example: Orbitals

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



Practice: Identify the Orbital

Identify the orbital shown below,



a) $2p_y$

b) $3d_{yz}$

c) $3p_x$

d) $3d_{xy}$

e) $3d_{x2-y2}$

4.3 Radial and Angulor Nodes

4.3.1

Angular and Radial Nodes

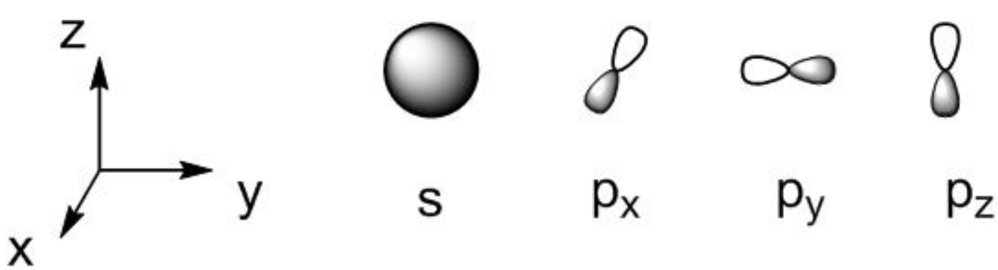
$$\boxed{Total\ number\ of\ nodes\ \ = n - 1}$$

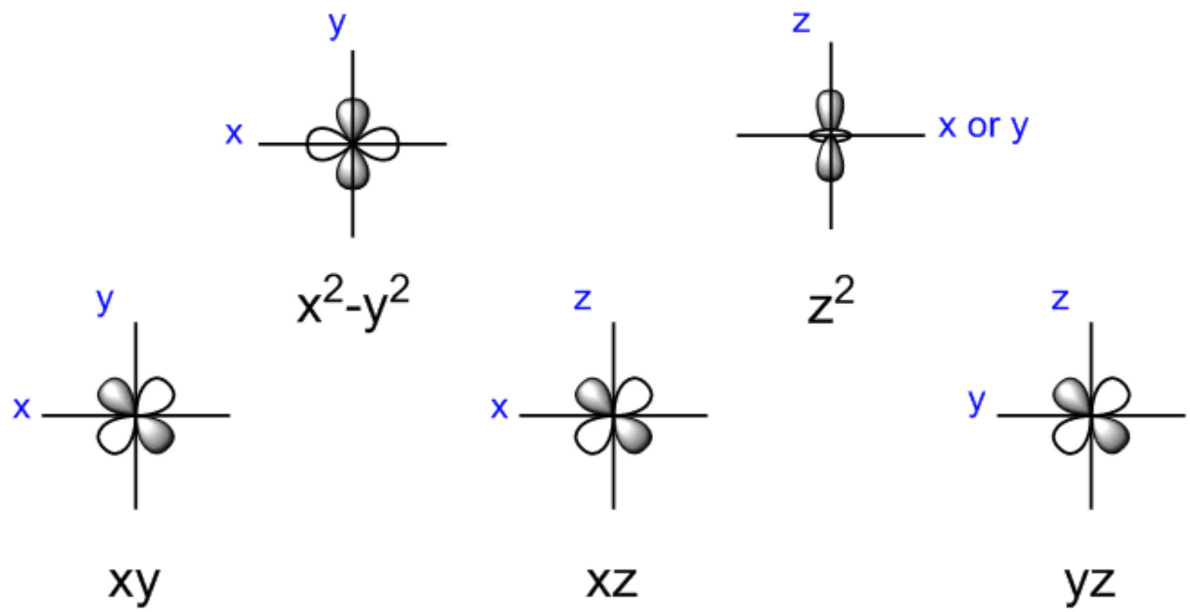
$$\boxed{Number\ of\ angular\ nodes = l}$$

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

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

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

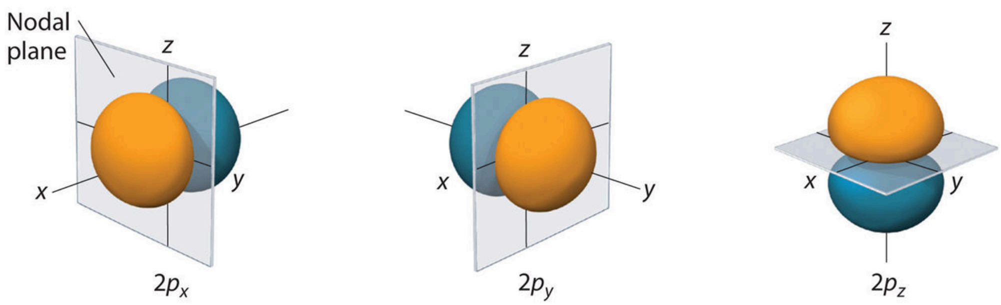


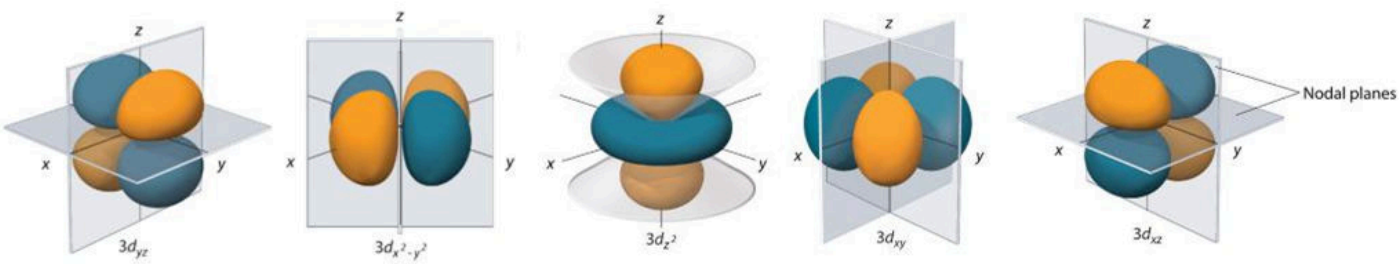


y
x
z
x or y
x^2-y^2
z^2
y
x
z
x
z
y
xy
xz
yz

Another View of the Nodes

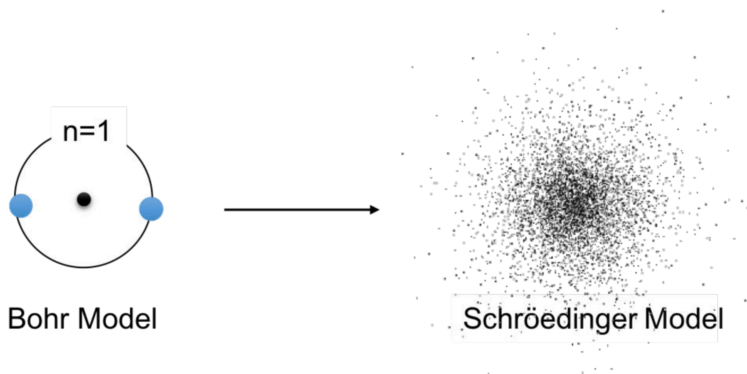
- **Nodes** are regions in space where the probability of finding an electron is zero.
- In p orbitals and d orbitals, we have **planar nodes** between the lobes of electron densities.
- We will take a look at s orbitals and their nodes soon...



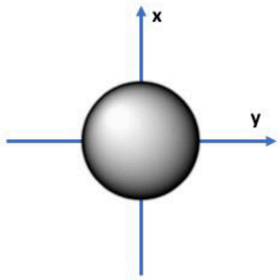


Wave Properties of Electrons

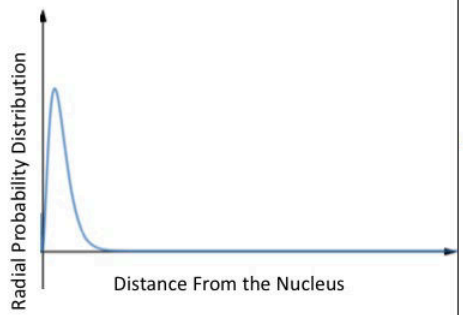
- Electrons have both particle and wave-like properties. Within an atom, electrons are standing waves, oscillating around the nucleus. As with all waves, there are nodes and phases.
- **Orbital:** mathematically derived region of probability in 3D space where an electron“may be” found



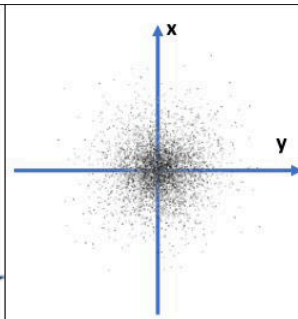
Representations of Atomic Orbitals



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



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



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

4.3.3

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.

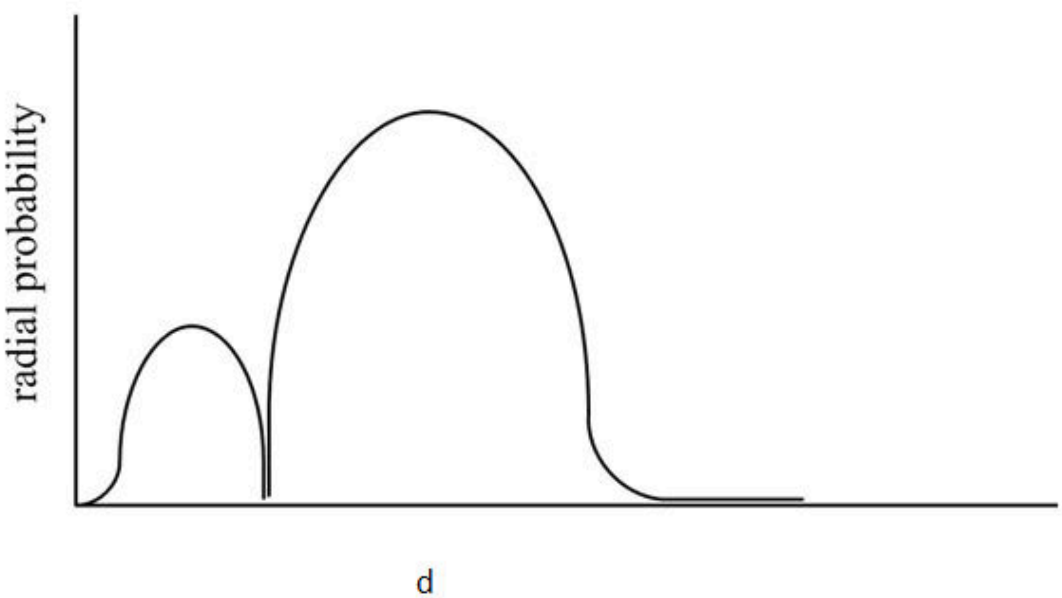
WIZE CONCEPT

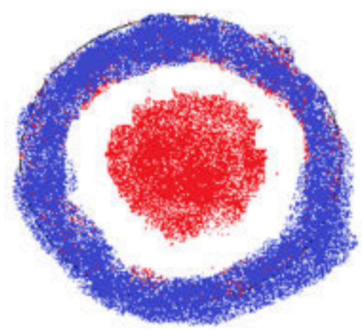
→ Each **radial node** will be indicated by a **zero point on a radial probability graph**

→ **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!**

4.3.4

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

4.3.5

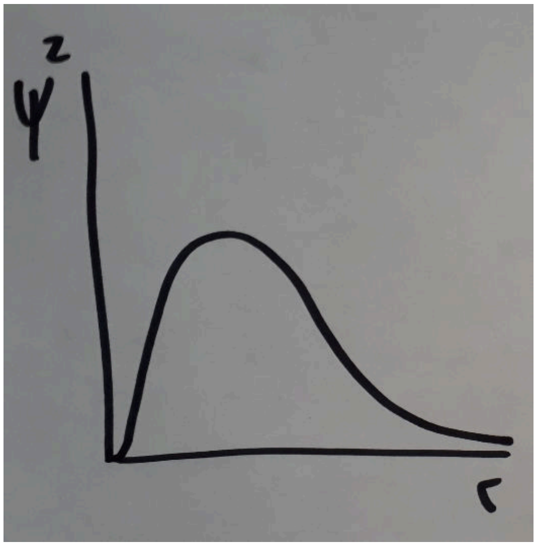
Example: Sketching Radial Probability

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

4.3.6

Example: Identifying Orbitals from Probability Plots

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



a) $3d_{z2}$

b)1s

c) $4p_x$

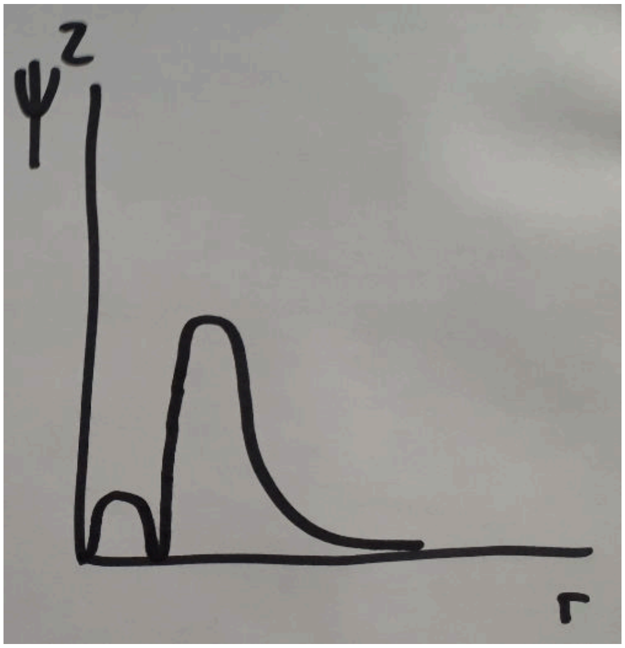
d) 3s

e) $2p_y$

4.3.7

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
- $3p_x$

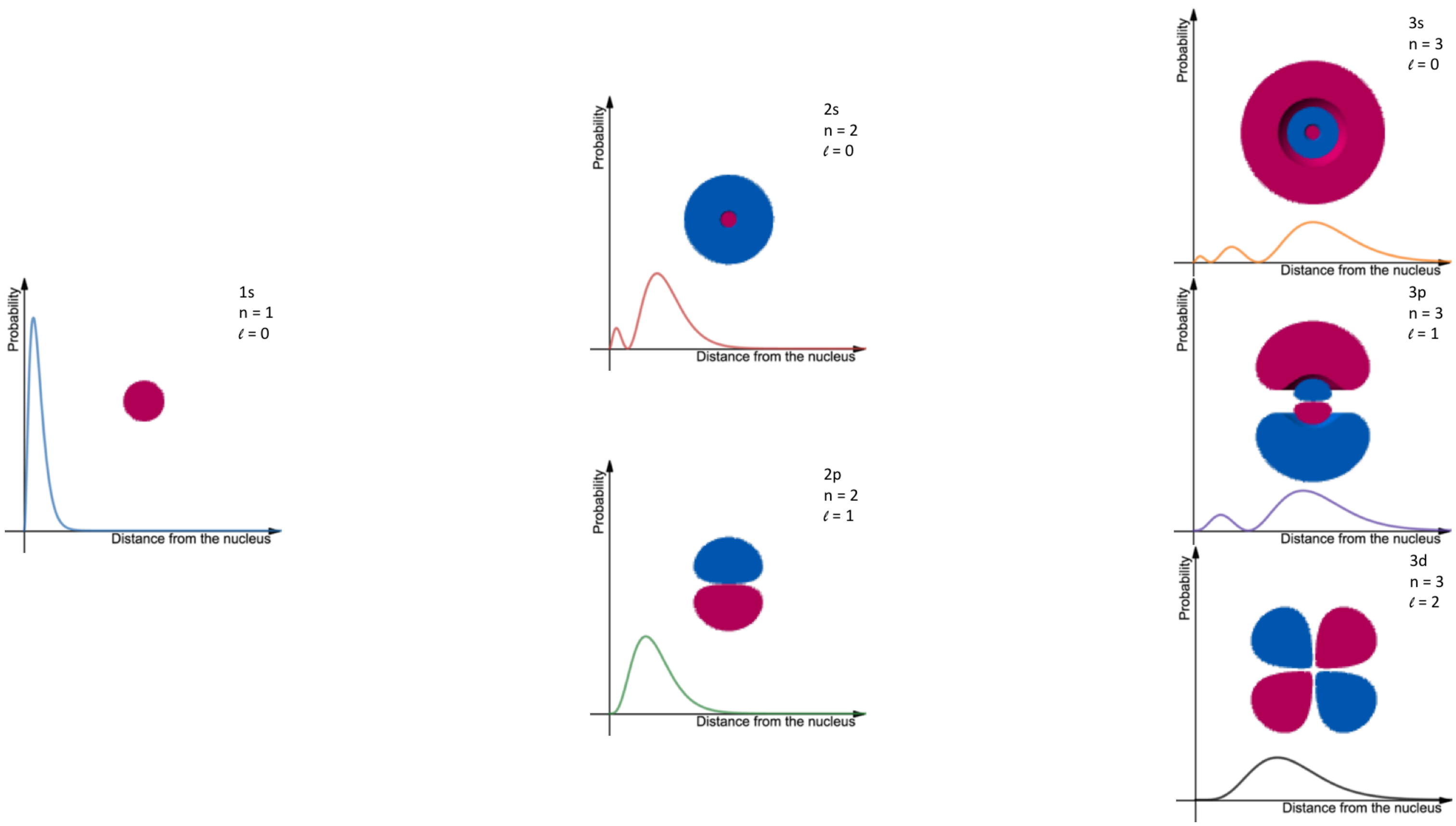
4.3.8

Nodes

- Nodes are regions in space where the probability of finding an electron is zero. There are *n-1* total nodes in an orbital.
  - In the first electron shell, n = 1. The 1s orbital has no nodes.
  - In the second electron shell, n = 2. The 2s and 2p orbitals have one node
  - In the third electron shell, n = 3. The 3s, 3p, and 3d orbitals have two nodes, etc.

Types of nodes

- **Angular nodes** are typically planes. The quantum number ℓ dictates the number of angular nodes in an orbital. They **do not** appear in radial probability diagrams
- **Radial nodes** are spheres that occur as the quantum number *n* increases. We can find the number of radial nodes from the quantum number *n* and the quantum number ℓ. There are n-ℓ-1 radial nodes in an orbital. They are shown in radial probability diagrams.



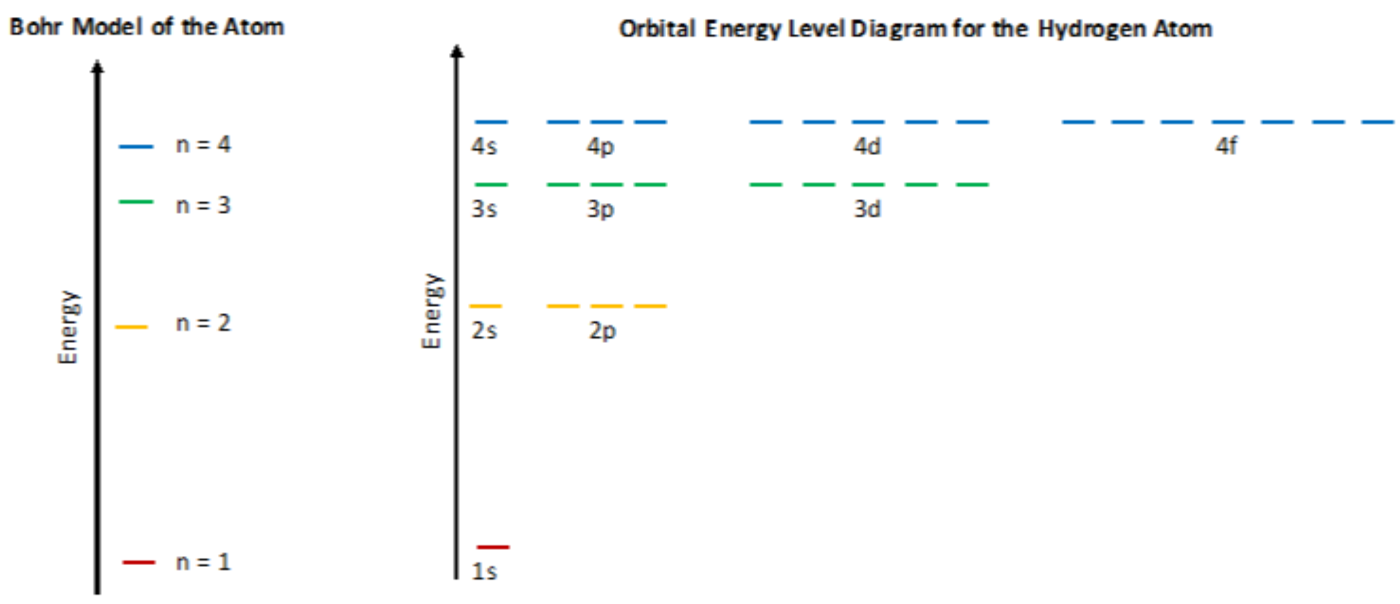
4.4 Orbital Filling Diagrams

4.4.1

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.**

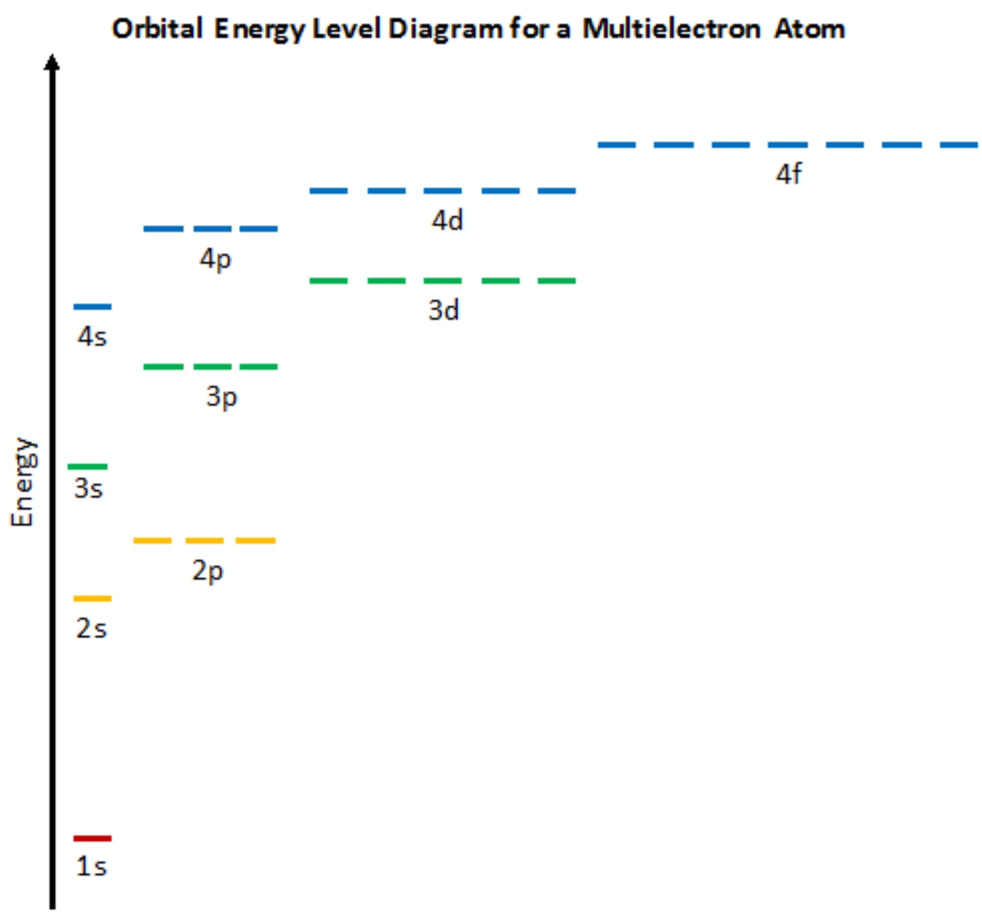


WIZE TIP

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

Multi-Electron Species

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



WIZE CONCEPT

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

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

Example:

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

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

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

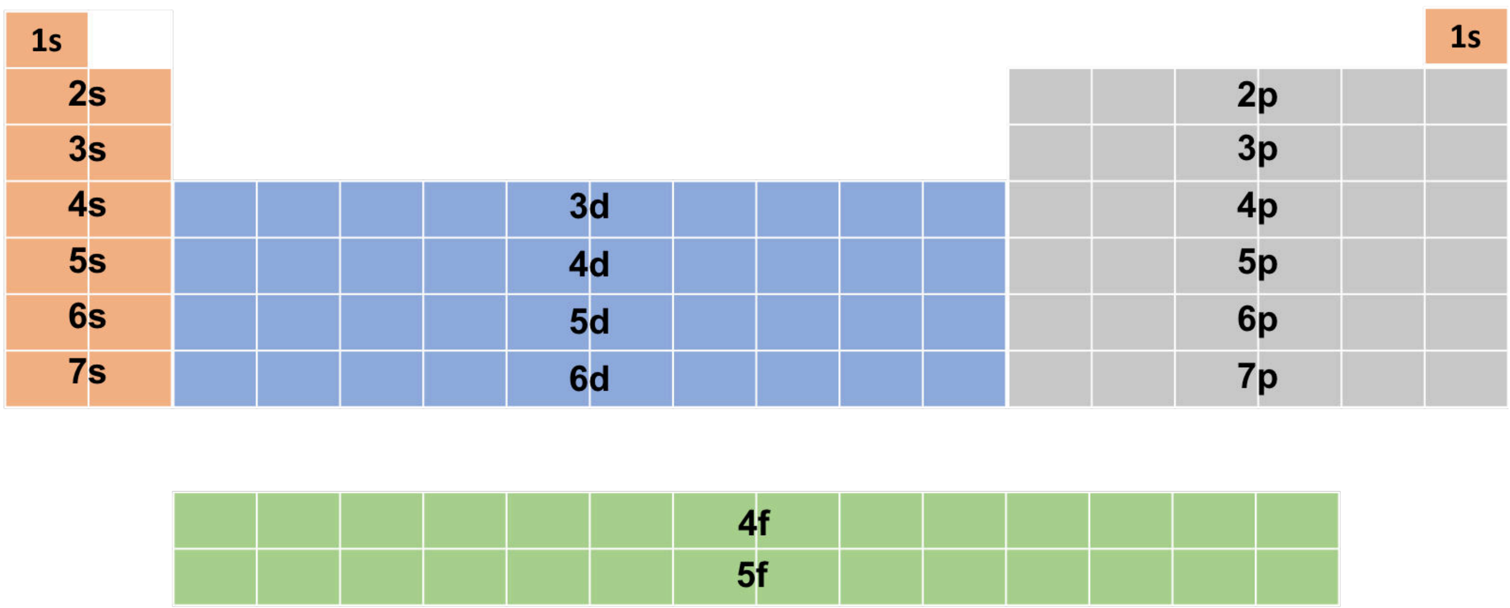
- We already found that total number of orbitals for this shell (_____)
- What is the maximum number of electrons we could have in each orbital? _____
- Therefore the maximum number of electrons possible when n=2 is _____

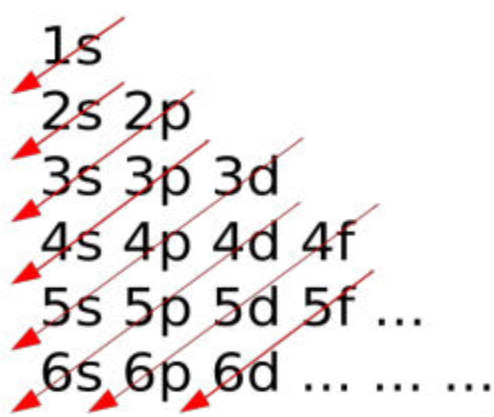
4.4.2

Rules for Orbital Filling

1) Aufbau Principle

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





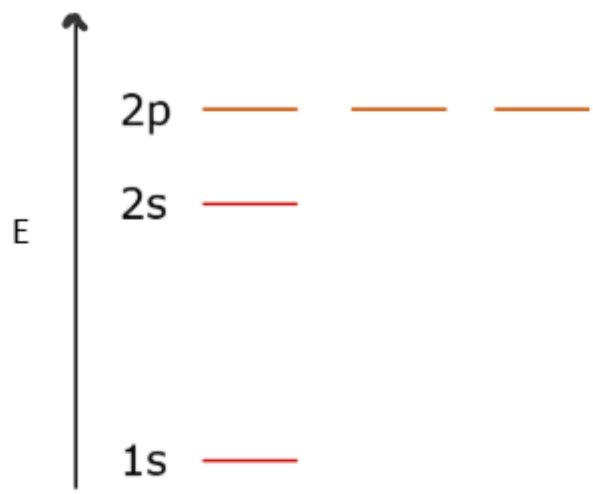
> **WATCH OUT!**
>
> **Memorize the order of orbital filling (or just familiarize yourself enough with the periodic table to know the order!):** 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p...

2) Hund's Rule

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

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

Example: Fill out the following orbital diagram for C



3) Pauli Exclusion Principle

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

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

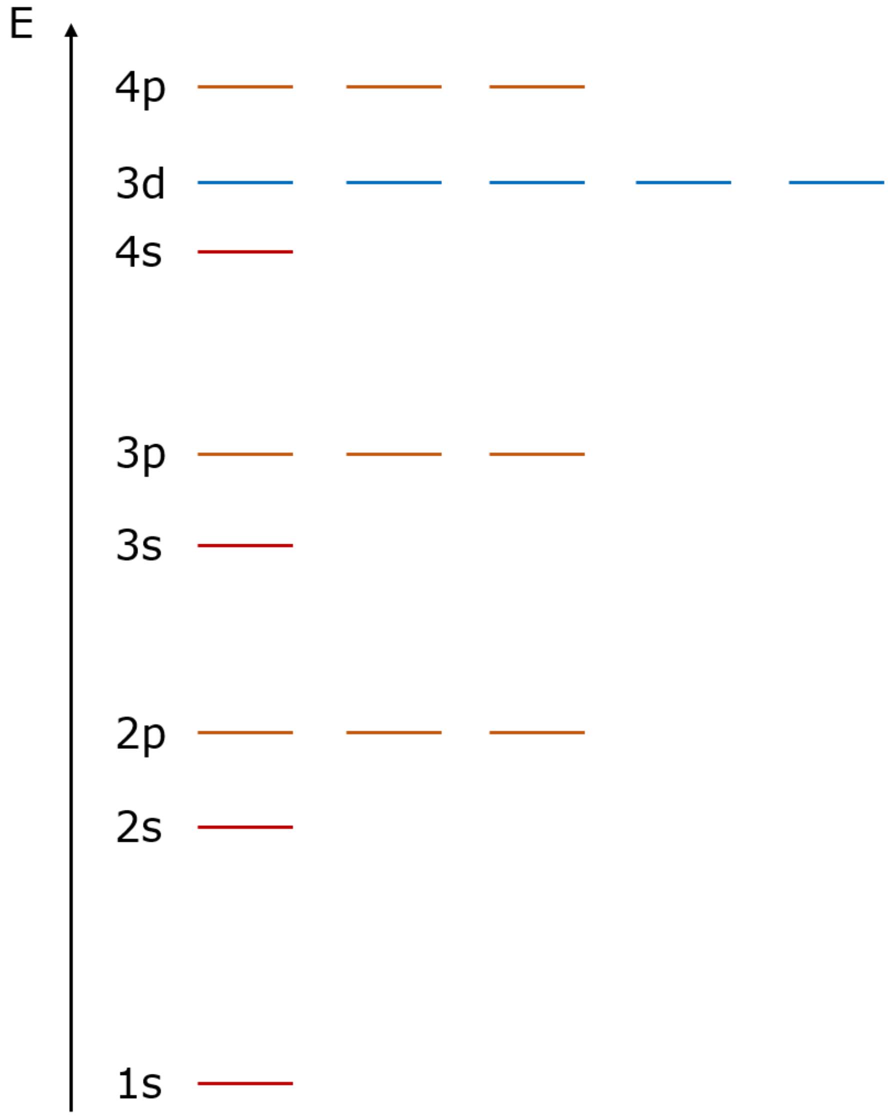
4.4.3

Example: Orbital Filling Diagrams

Draw the orbital diagram for oxygen.

Example: Orbital Filling Diagrams

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

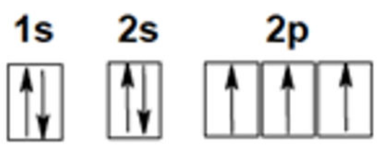


4.4.5

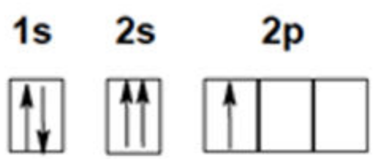
Practice: Orbital Filling Diagrams

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

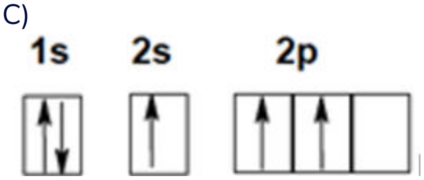
A)



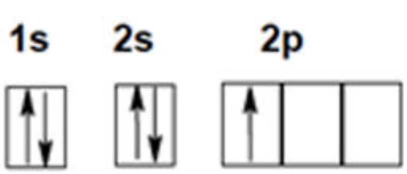
B)



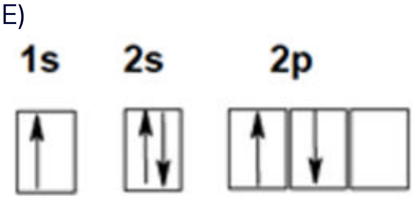
C)



D)



E)



4.4.6

Practice: Number of Orbitals

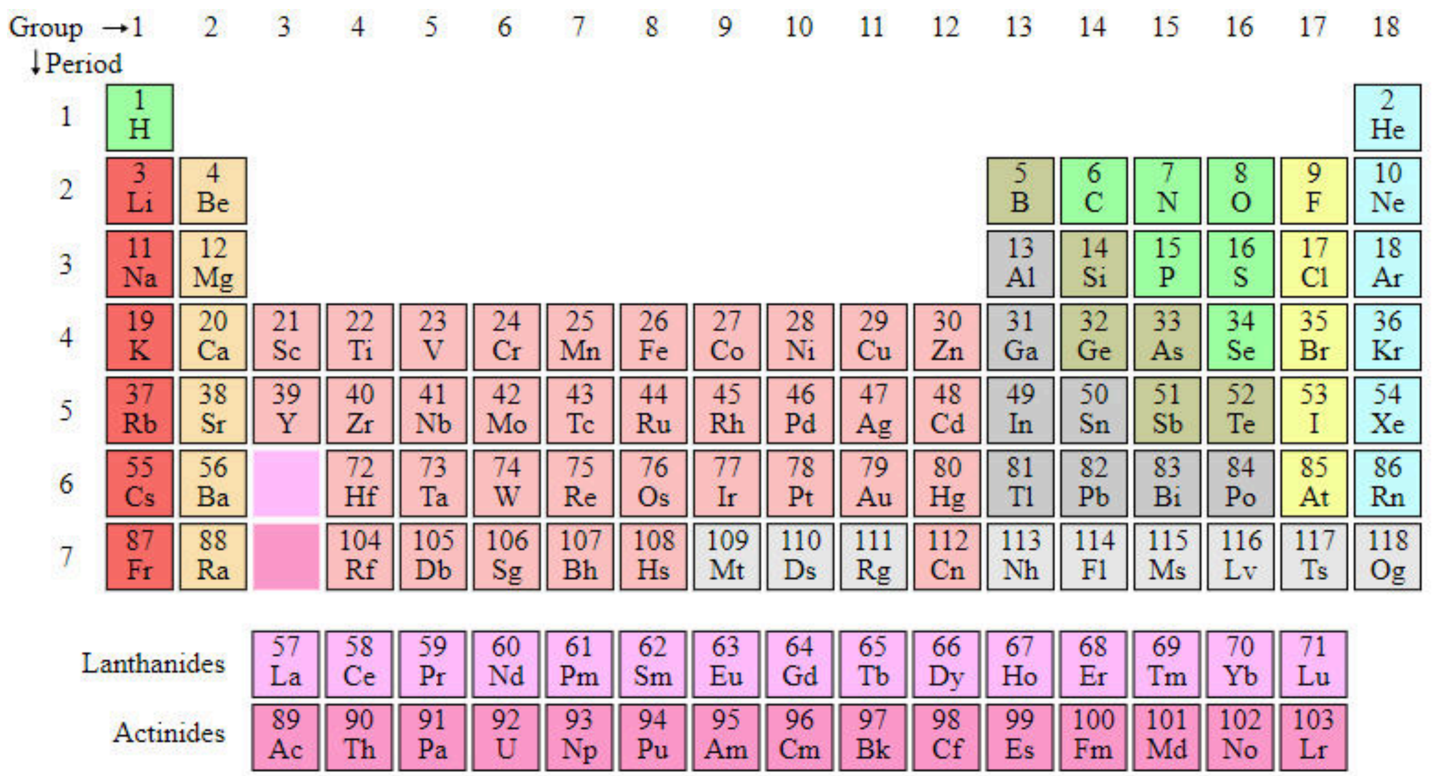
How many different atomic orbitals exist where n = 3?

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

4.5 Electronic Configuration for Atoms

4.5.1

Electron Configurations of Atoms



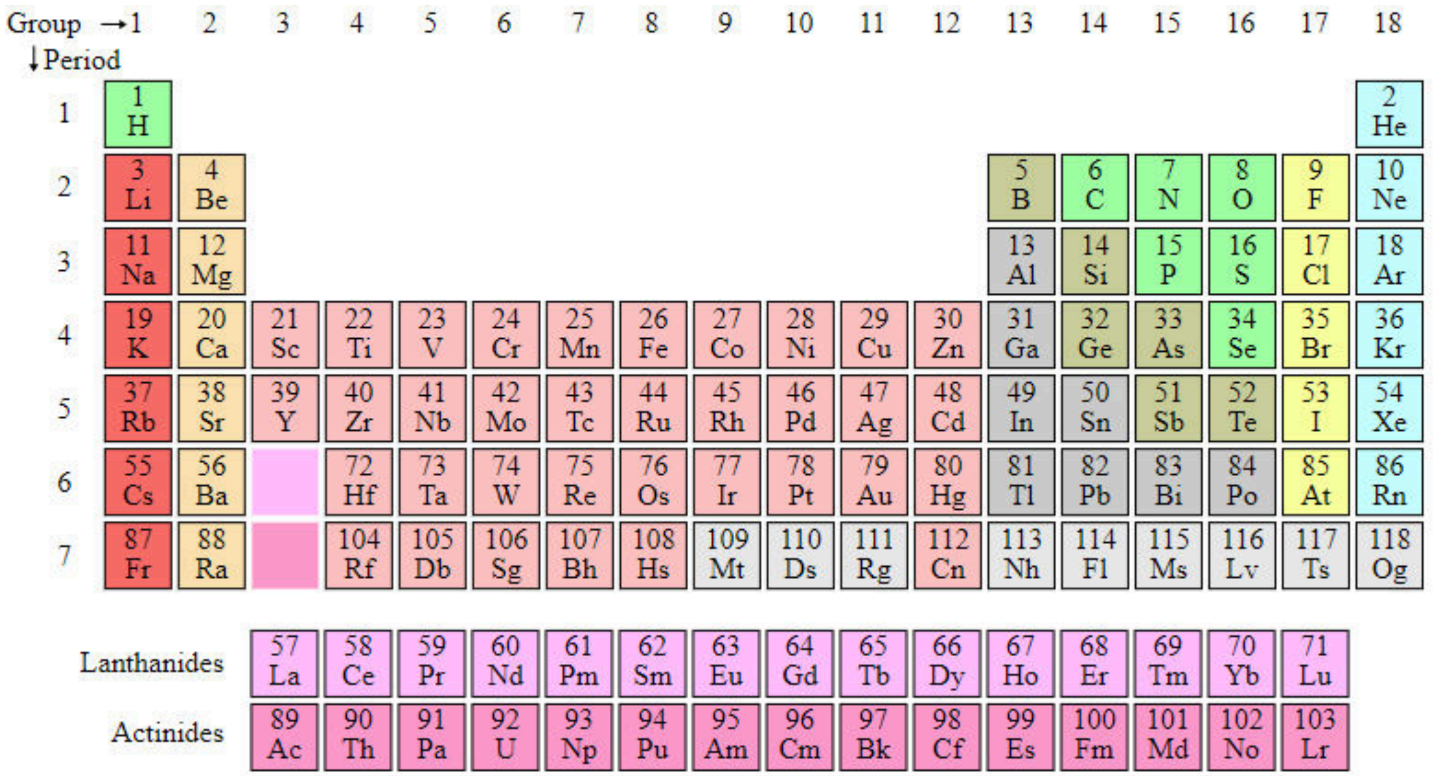
- lable blocks
- describe the n number for each block

Simple Electron Configurations

Li:

C:

Ne:



Shorthand Notations

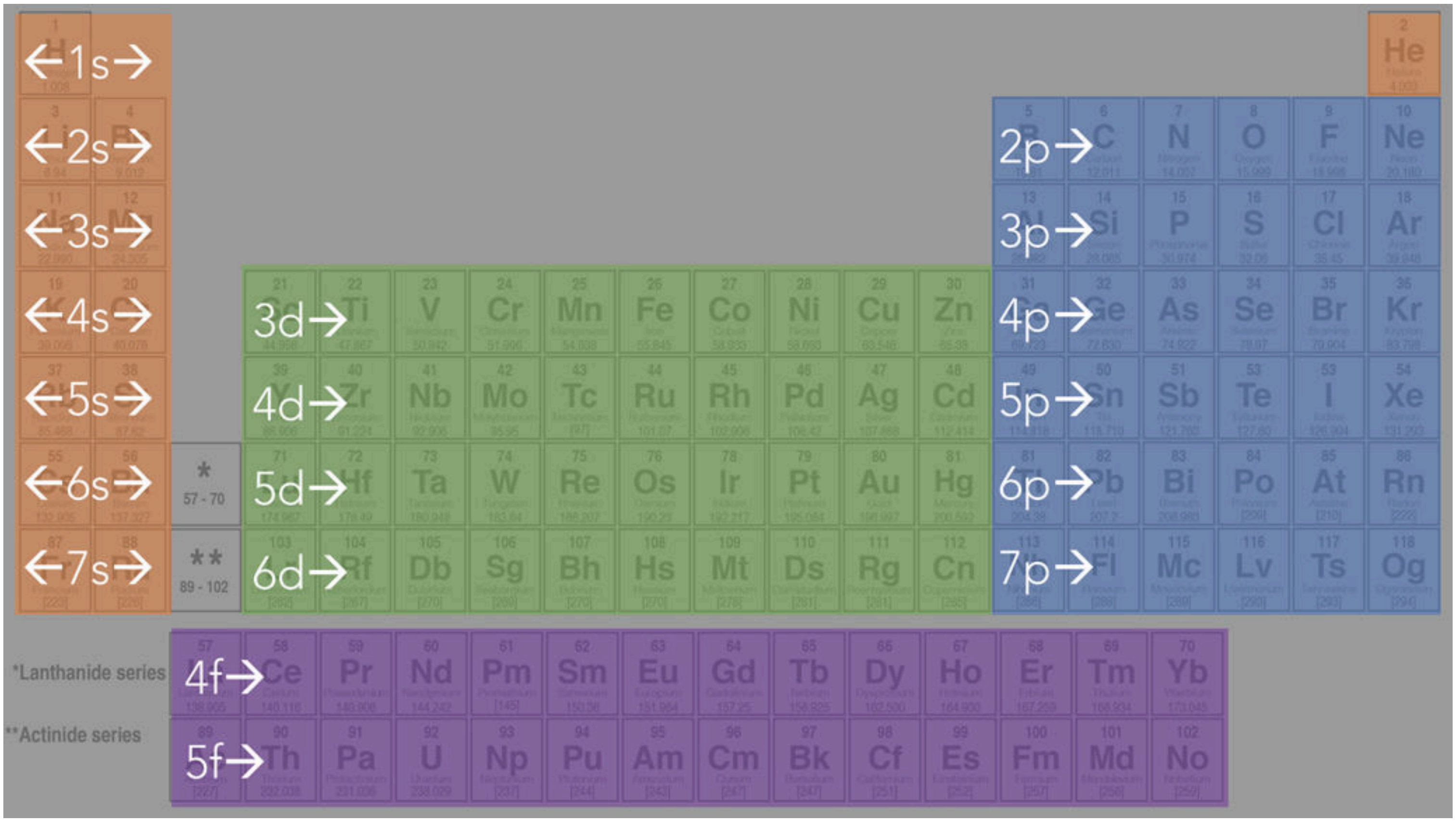
Write the name of the **previous noble gas and then fill in the rest of the electron configuration.**

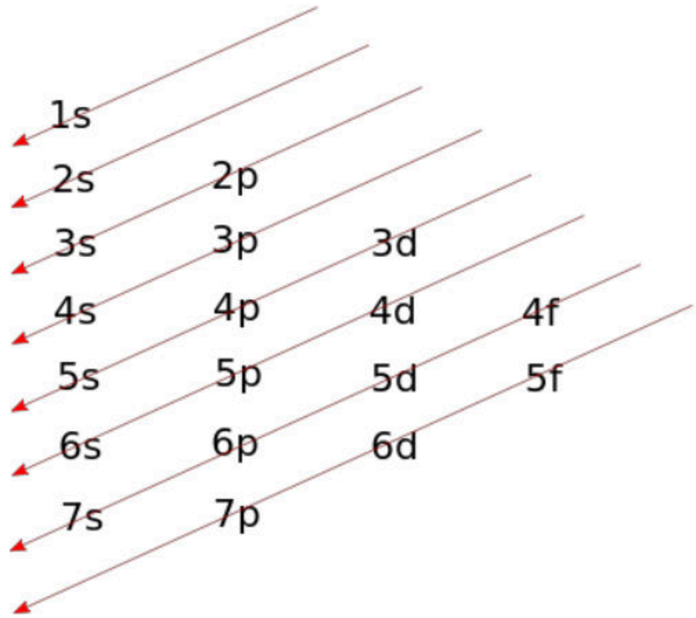
O:

Fe:

Ge:

Electron Configurations Cheatsheet




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

4.5.3

Electron Configuration Exceptions

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

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

WIZE TIP

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

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:

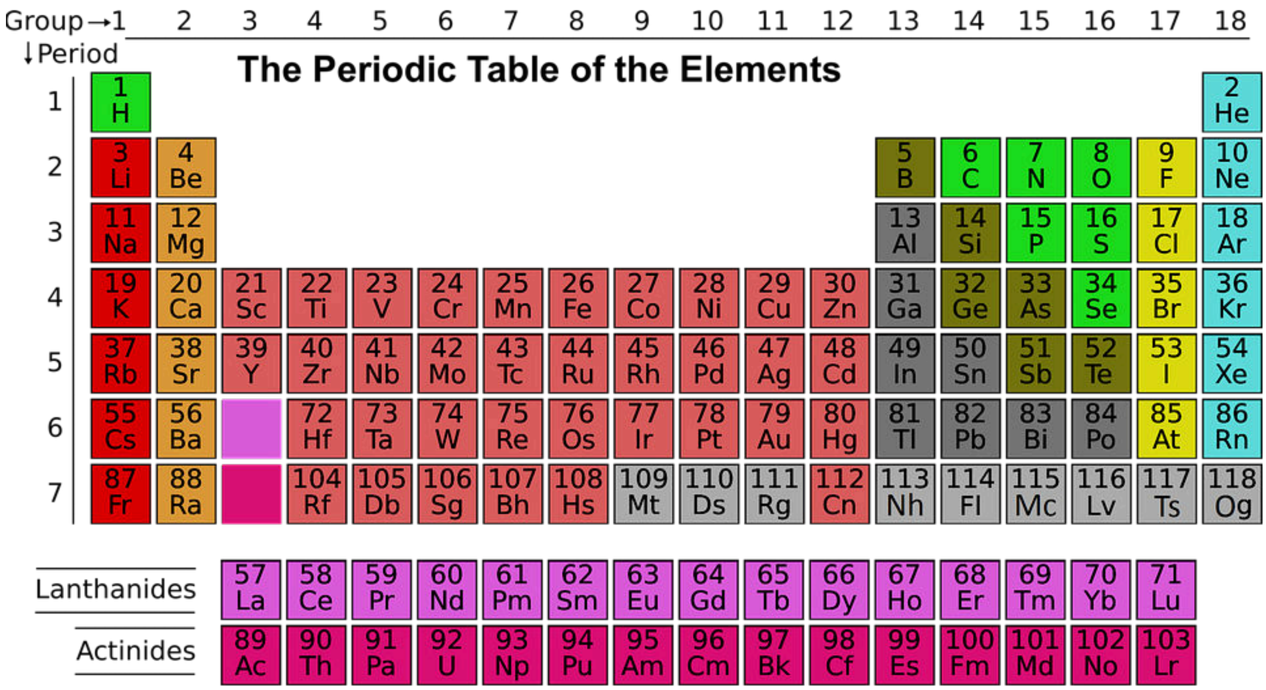
Zr:

The Periodic Table of the Elements

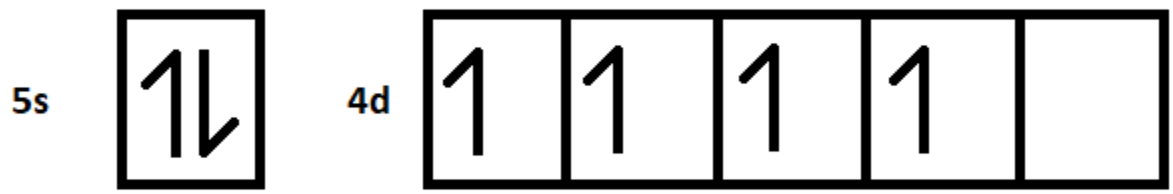
Group → ↓Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
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 Mc	116 Lv	117 Ts	118 Og
Lanthanides			57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu	
Actinides			89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr	

Example: Condensed Electron Configuration

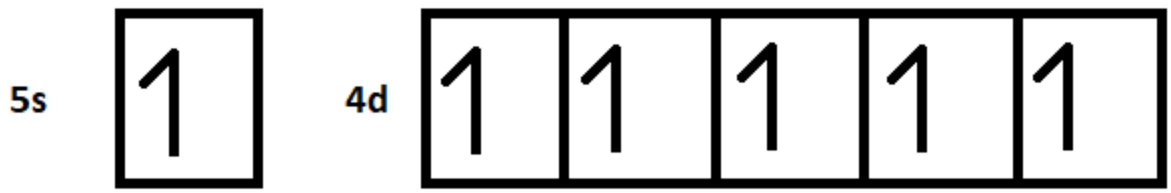
What is the condensed electronic configuration for Mo?



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



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



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

4.5.6

Practice: Identify the Neutral Element

What neutral element is represented in the example below?

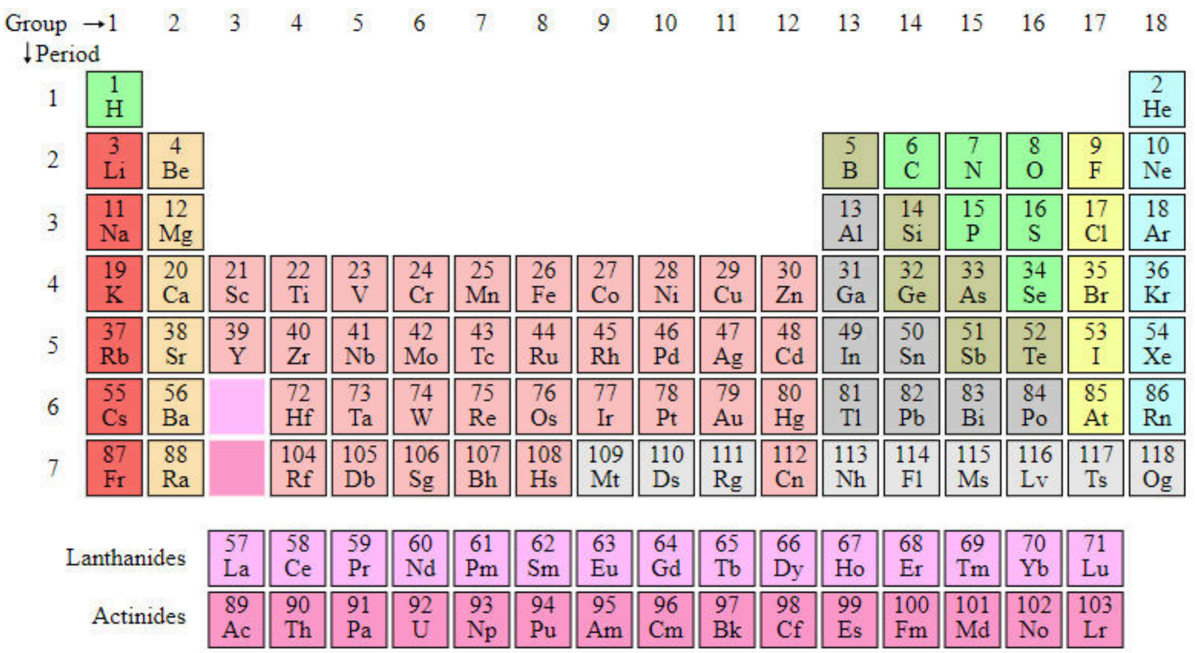
$1s^2 2s^2 2p^6 3s^2 3p^5$ became $[Ne]3s^2 3p^5$

- Al
- Ne
- P
- Cl

4.5.7

Practice: Electron Configuration

The ground state electronic configuration for At is



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

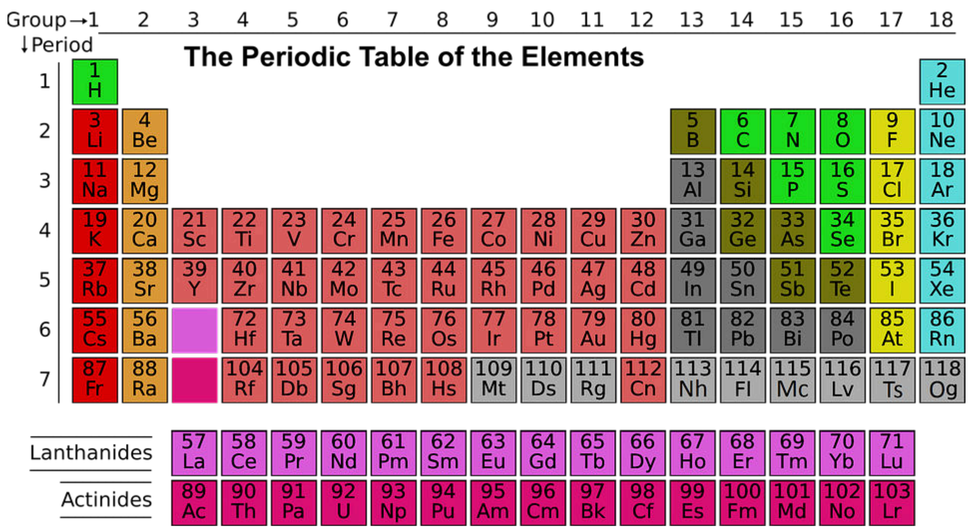
B) $[Xe]6s^25d^{10}6p^5$

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

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

Practice: Impossible Electron Configuration

Which of the following electronic configurations is not possible?



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

4.6 Electron Configurations for Ions

4.6.1

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: **2**s has a higher energy than **1**s

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

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

What is the electron configuration for F^+?

Solution available online

WIZE TIP

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

Example: F^- is isoelectronic with Ne!

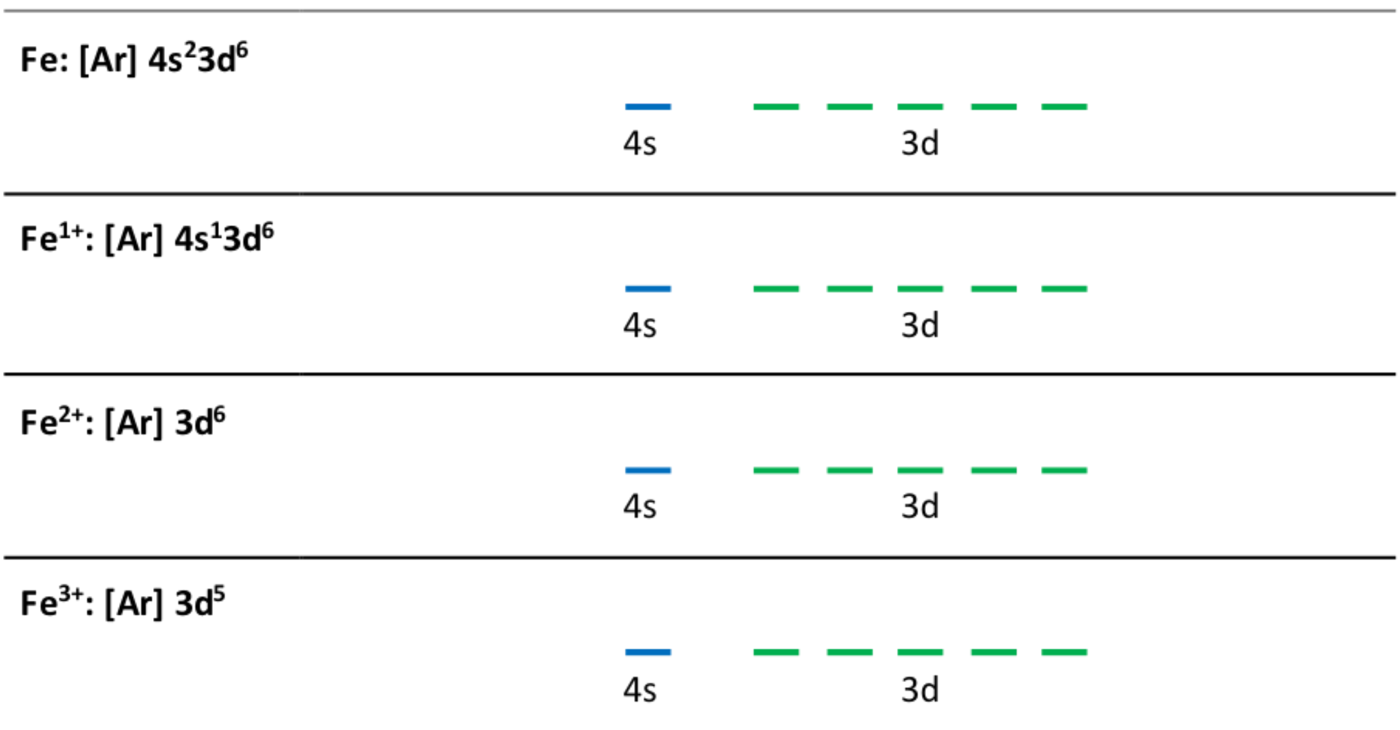
4.6.2

Electron Configuration of Ions of Transition Elements

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

WATCH OUT!

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



WIZE TIP

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

This will help you avoid mistakes on the exam!

4.6.3

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 → / ↓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) N^+ and C

ii) C and B^+

iii) Cl^- and Ar

4.6.4

Example: Writing Electron Configurations of Ions

Write the electronic configuration for the following species:

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

a) N^+

b) Ni^{2+}

c) Ge^-

4.6.5

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

4.7 Diamagnetic vs Paramagnetic Electron Configurations

4.7.1

Diamagnetic vs Paramagnetic

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

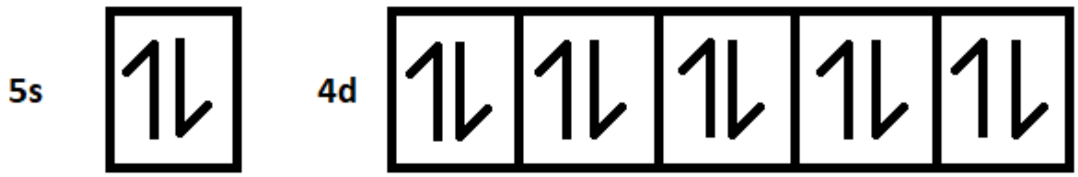
Diamagnetic - all electrons are **paired**

Paramagnetic - all electrons are **not paired**

Are the following electron configurations diamagnetic or paramagnetic?

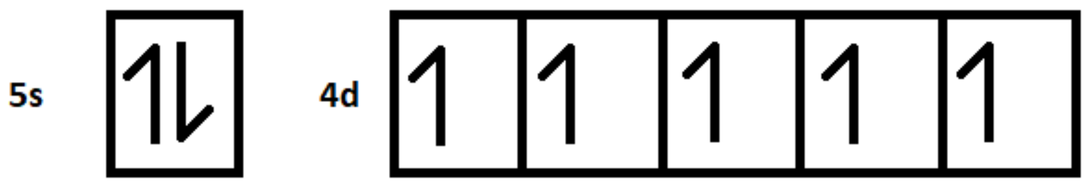
Cd

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



Tc

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



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

Consider the electron configuration for Na

Is it diamagnetic or paramagnetic? ______________

Consider the electron configuration for Si

Is it diamagnetic or paramagnetic? ______________

Consider the electron configuration for Be

Is it diamagnetic or paramagnetic? ______________

WIZE CONCEPT

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

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

4.7.2

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+}

c) Mn^{7+}

4.7.3

Practice: Diamagnetic

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

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	

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

4.8 Ground State vs Excited State Electron Configurations

4.8.1

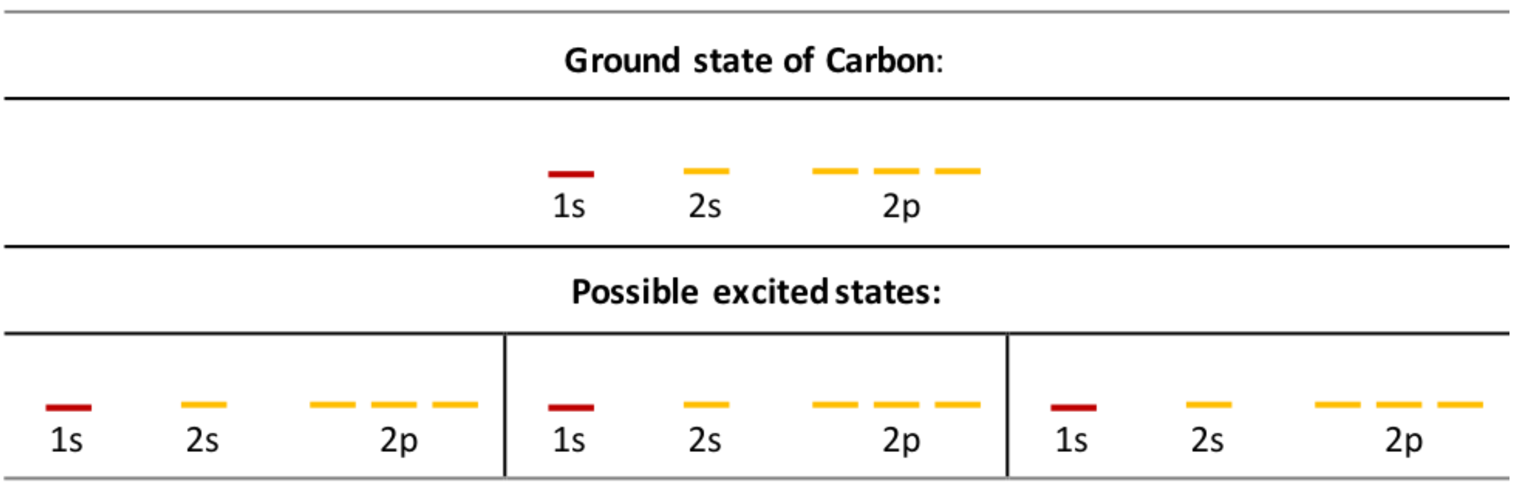
Ground State Vs Excited State Electron Configurations

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

WATCH OUT!

Do not confuse excited state configurations with those of ions.

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



WIZE TIP

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

Example: Excited State Electronic Configurations

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

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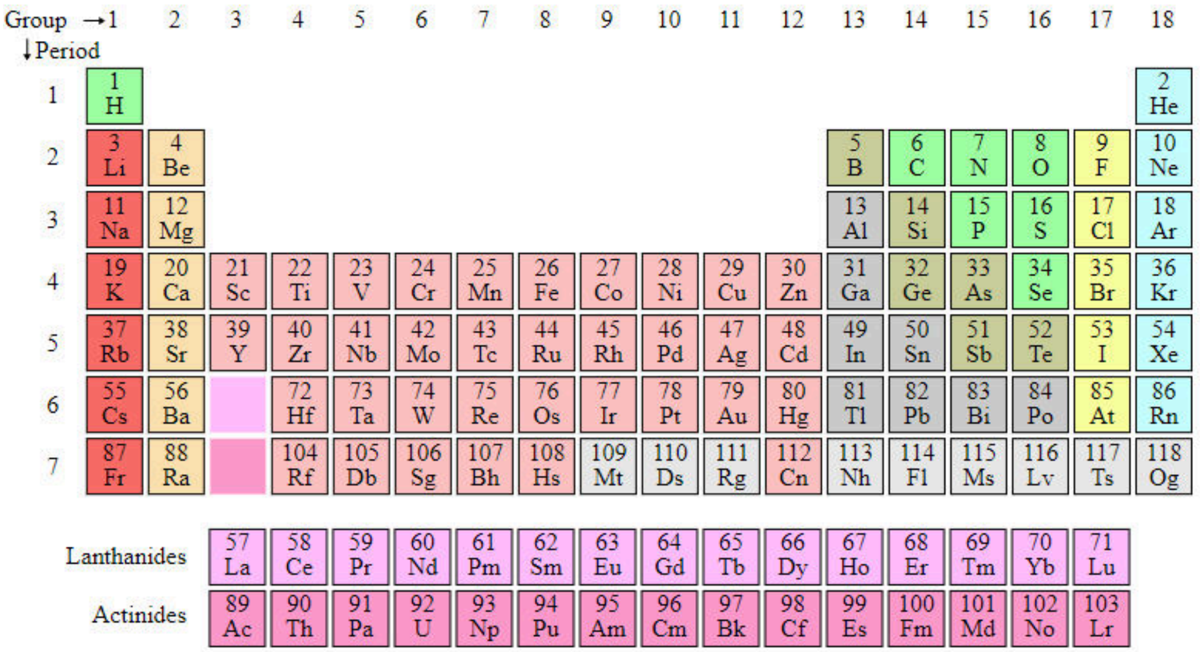
Lanthanides	57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
Actinides	89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

4.8.3

Example: Excited States

How many electrons would an excited atom of chlorine have?



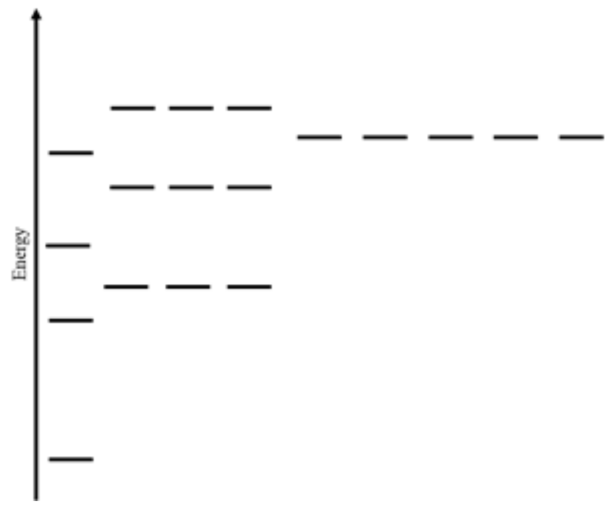
A. 6
B. 7
C. 16
D. 18
E. None of the above

4.8.4

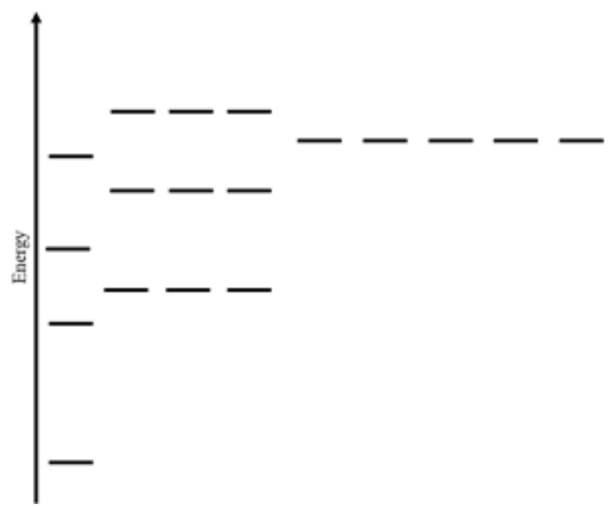
Practice: Excited State in an Energy Level Diagram

Consider an atom of the alkali metal Potassium.

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



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



4.8.5

Practice: Ground or Excited State

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

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

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