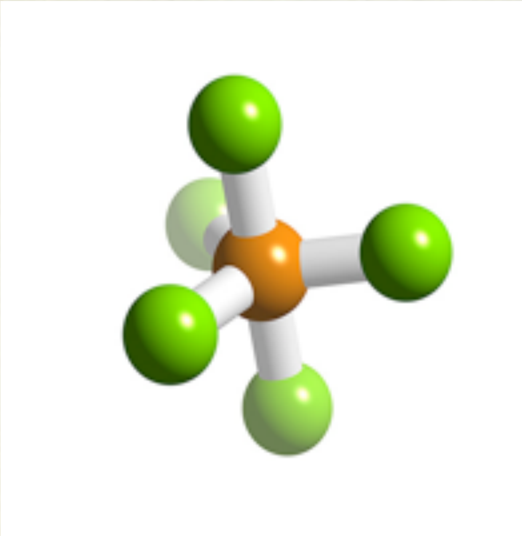
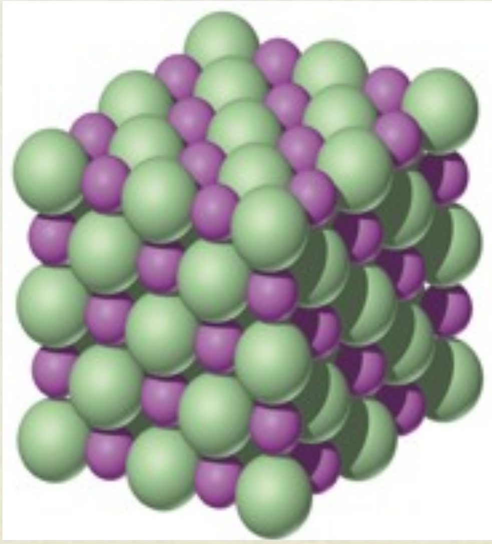
STRUCTURE & PROPERTIES OF MATTER

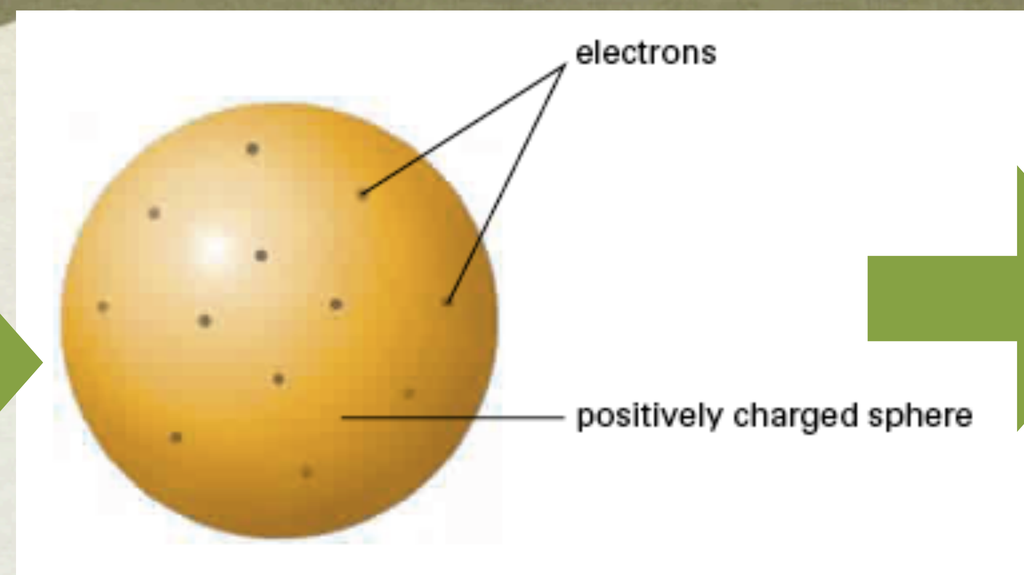


Chapter 3 & 4
in Chemistry 12

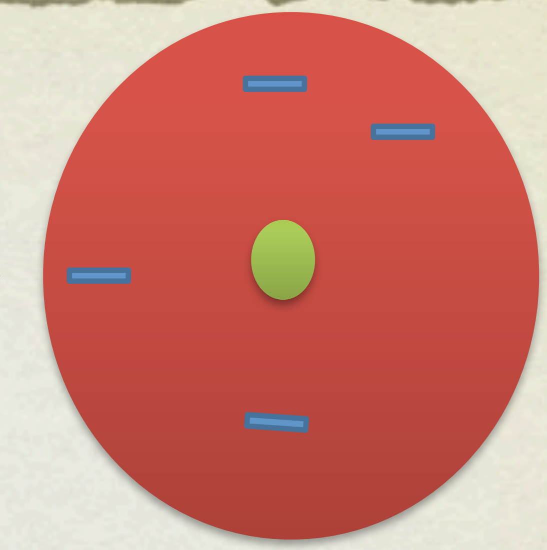


Recap!

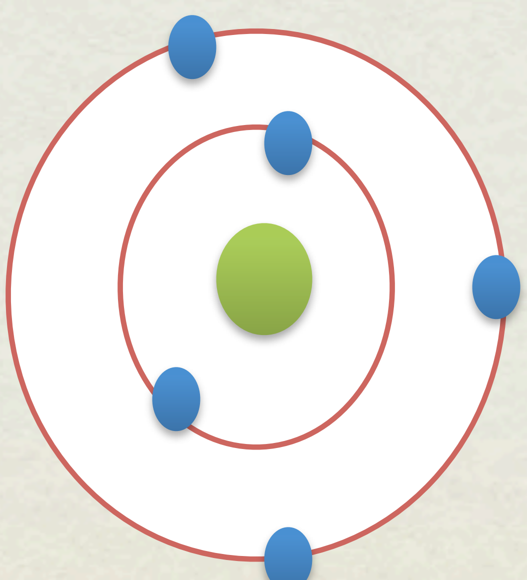
Dalton



Thomson

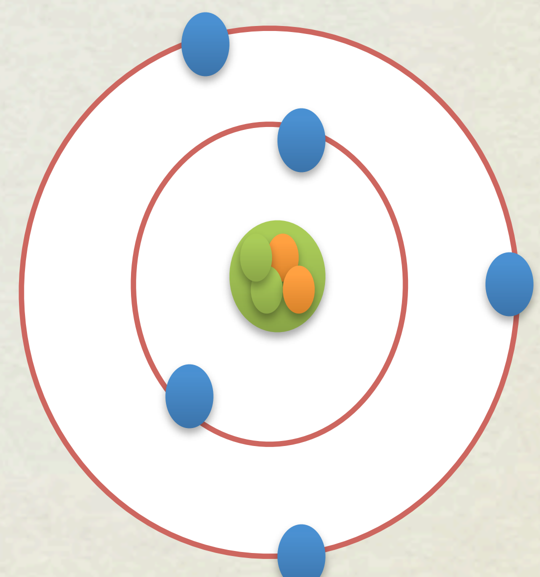


Rutherford



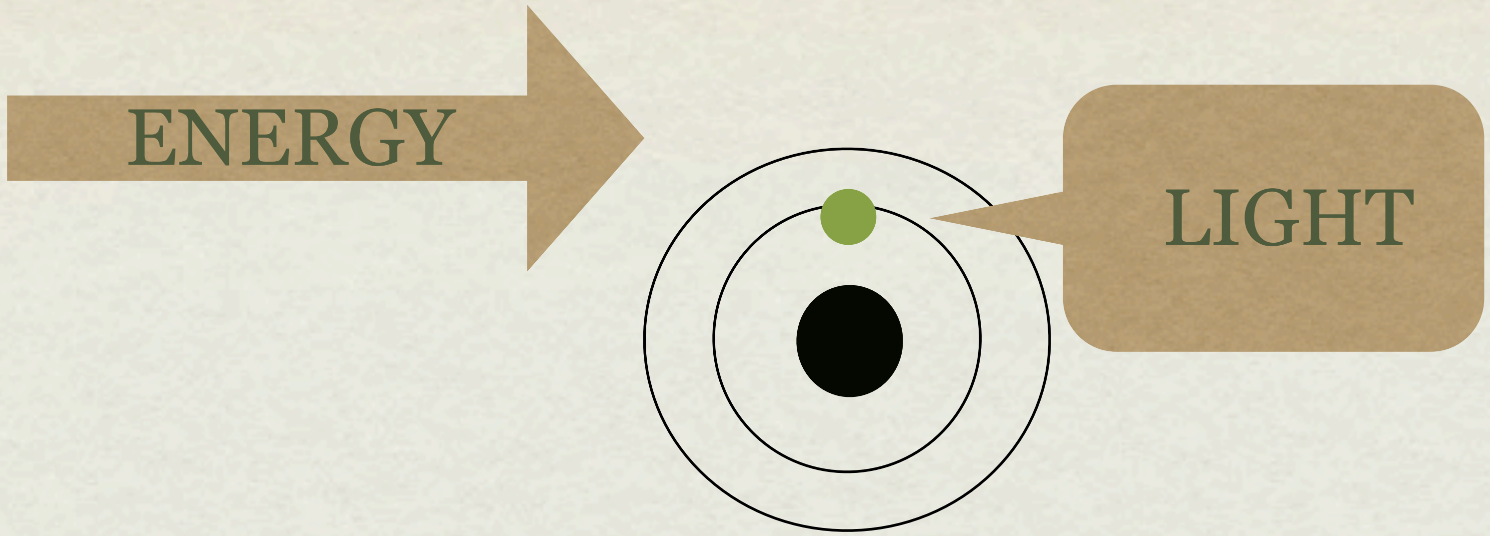
Bohr

+ Chadwick

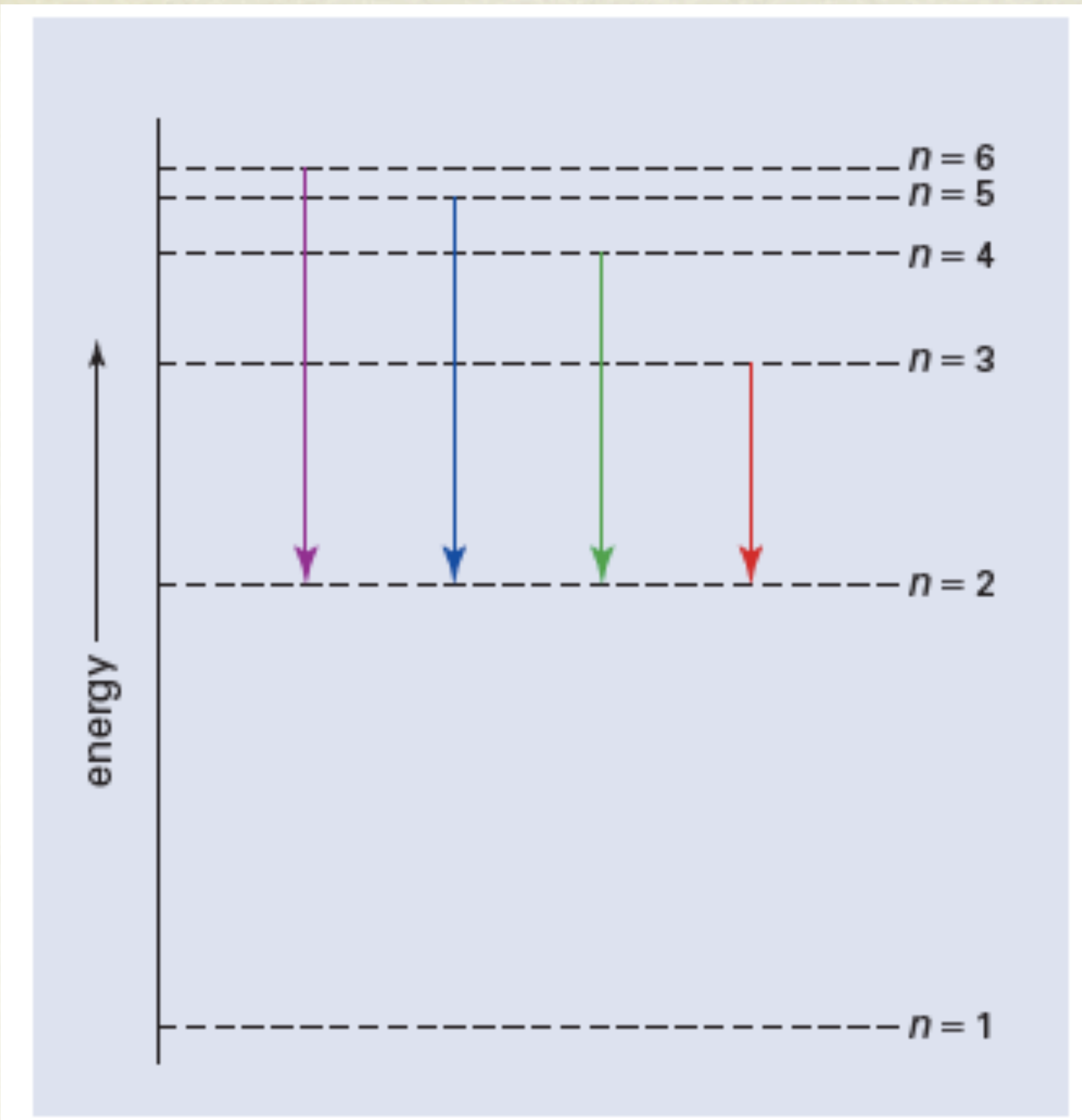


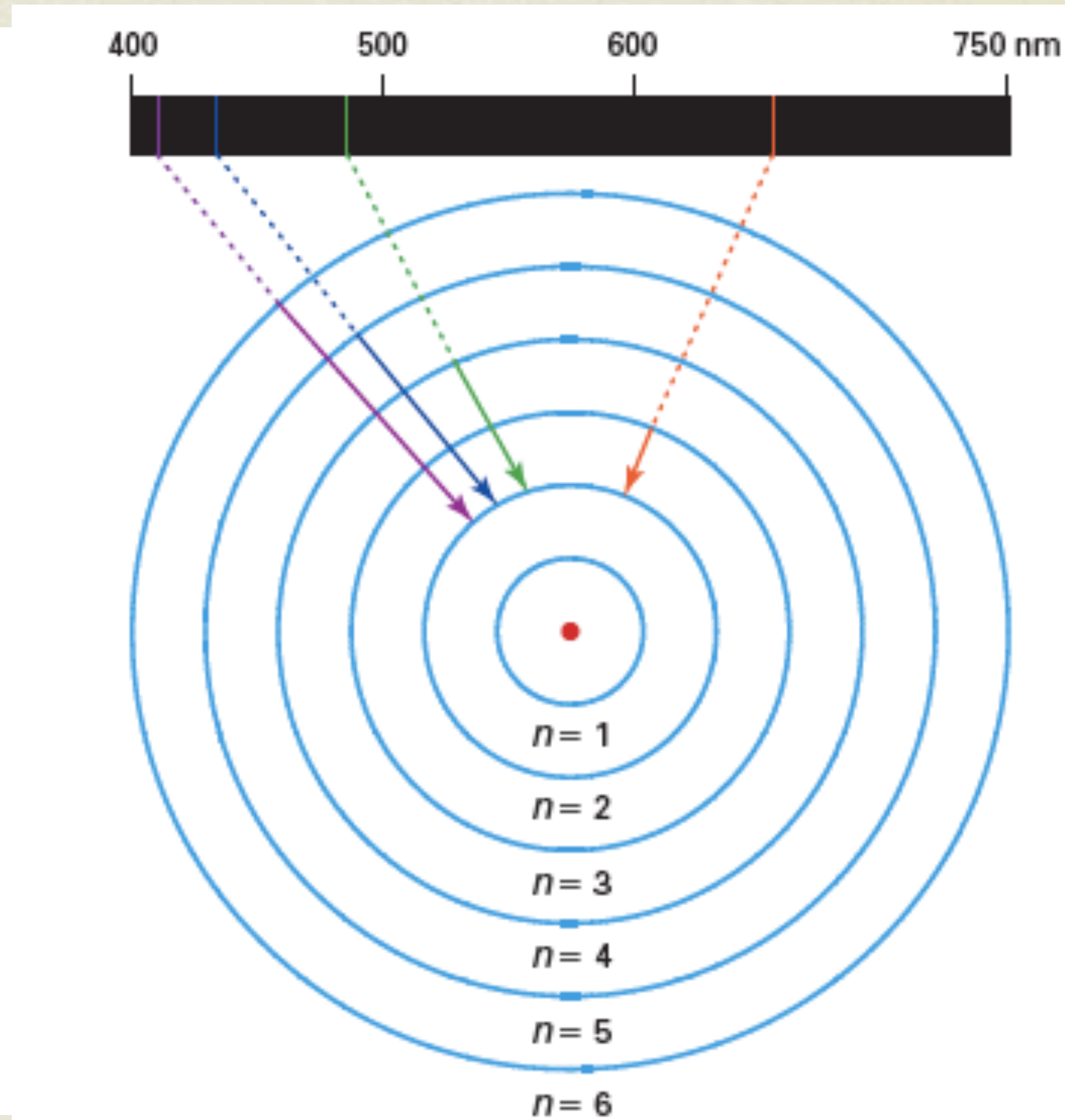
Bohr-Rutherford

Exciting Atoms



Spectra of Hydrogen



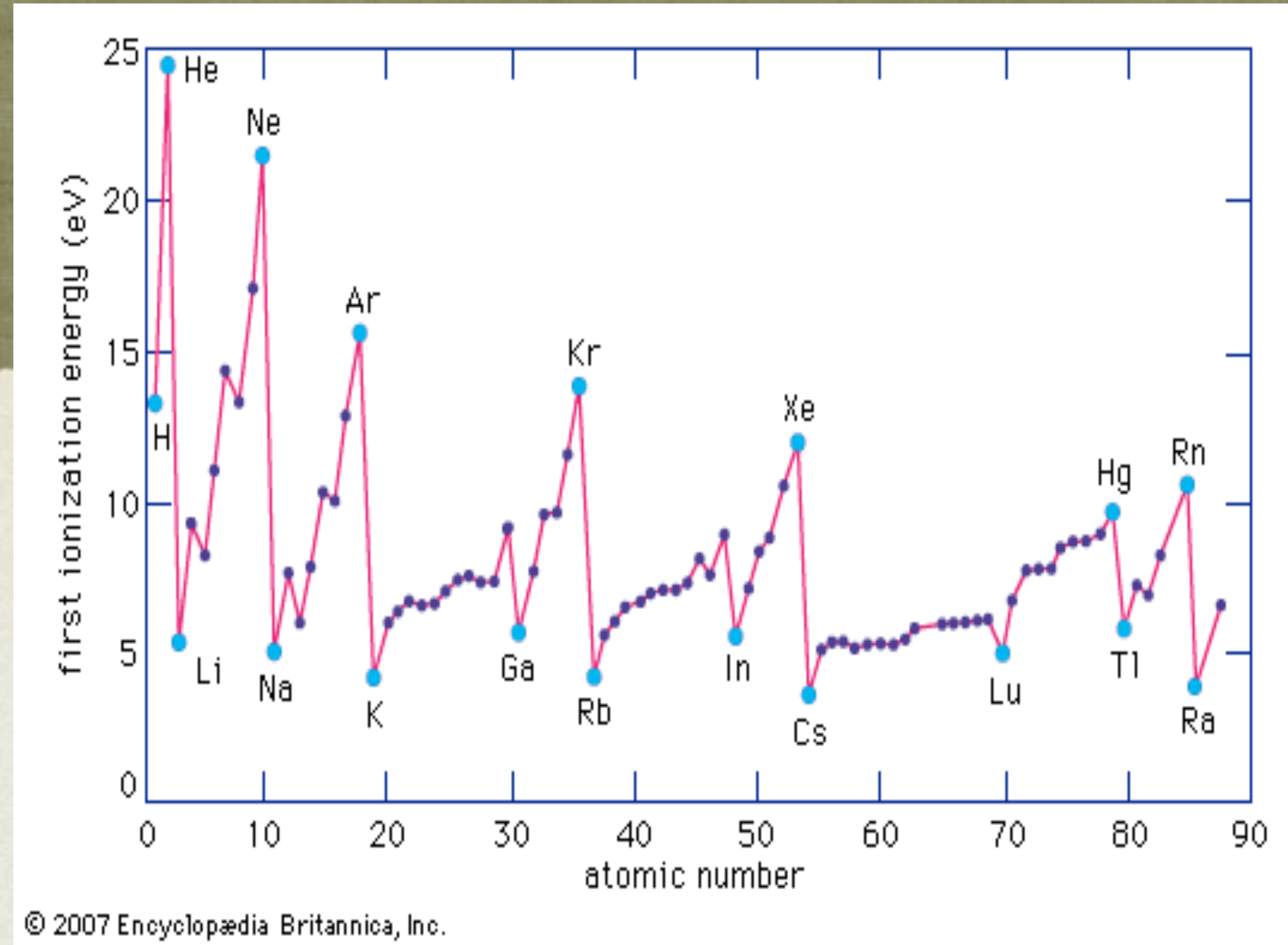


Quantum Mechanics

- We left off by saying Bohr's model only explained the electron arrangement of Hydrogen...
- A new model was needed.
- Because of observations made regarding ionization energy and line spectra, it was concluded that we needed to take into consideration more than just the energy level of an electron
- Enter - **QUANTUM MECHANICS!**

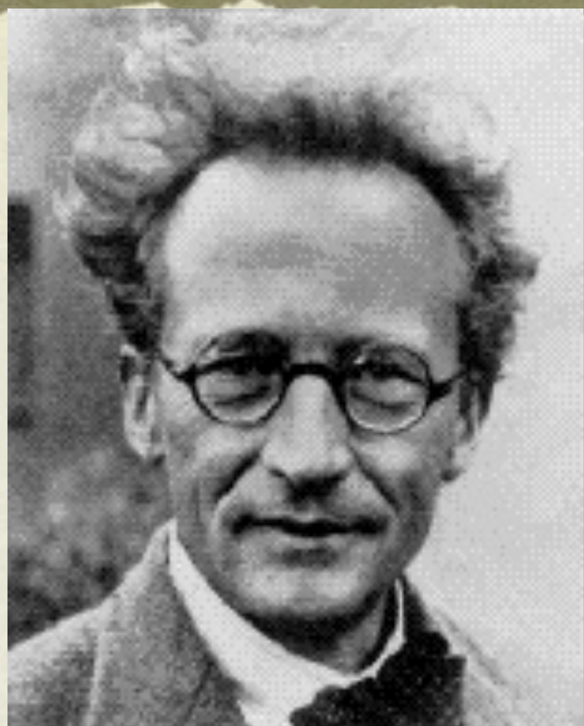
Ionization energy

- The first ionization energy of an element is the energy required to remove one electron from an atom
- We can see a pattern



- Why is there a difference between Li and Ne ? Draw a BR diagram of both.

Quantum Mechanics



E. Schrodinger

1887-1961

Schrodinger applied idea of electrons behaving as a wave to the problem of electrons in atoms.

He developed the **WAVE EQUATION**

Solution gives set of math expressions called WAVE FUNCTIONS, Ψ

The solutions to the Schrodinger wave equation describe the 3D shapes of the atomic orbitals where there is a high probability that electrons are located.

Heisenberg "Uncertainty Principle"

W. Heisenberg

1901-1976

- Electrons behave as both waves and particles.
- Heisenberg determined that it is impossible to know **BOTH** the exact position and momentum of an electron.
- He observed that on cannot simultaneously define the position and momentum (= m•v) of an electron.
- If we define the energy exactly of an electron precisely we must accept limitation that we do not know exact position.

QUANTUM NUMBERS

The shape, size, and energy of each orbital is a function of 4 quantum numbers which describe the location of an electron within an atom or ion

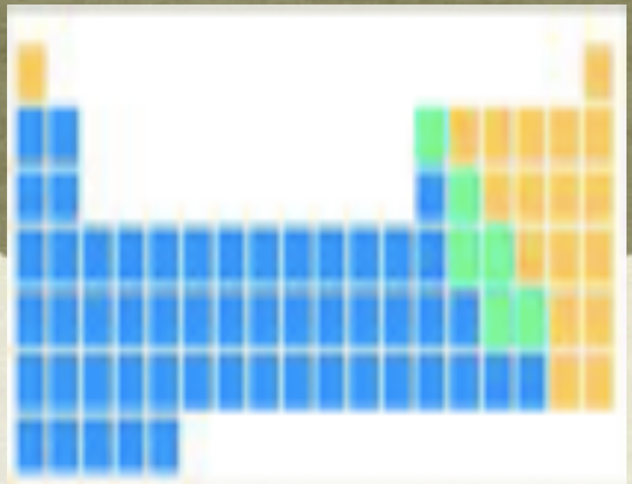
n (principal) --->

ℓ (orbital) --->

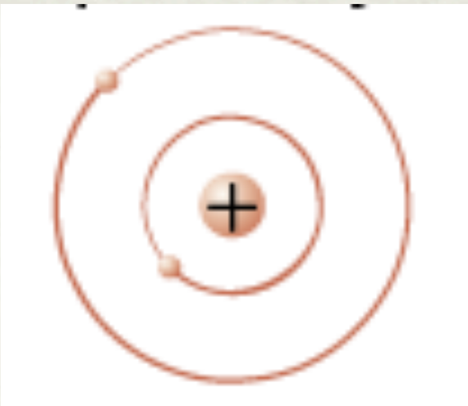
m_l (magnetic) --->

m_s (spin) --->

Principal Quantum Number, n



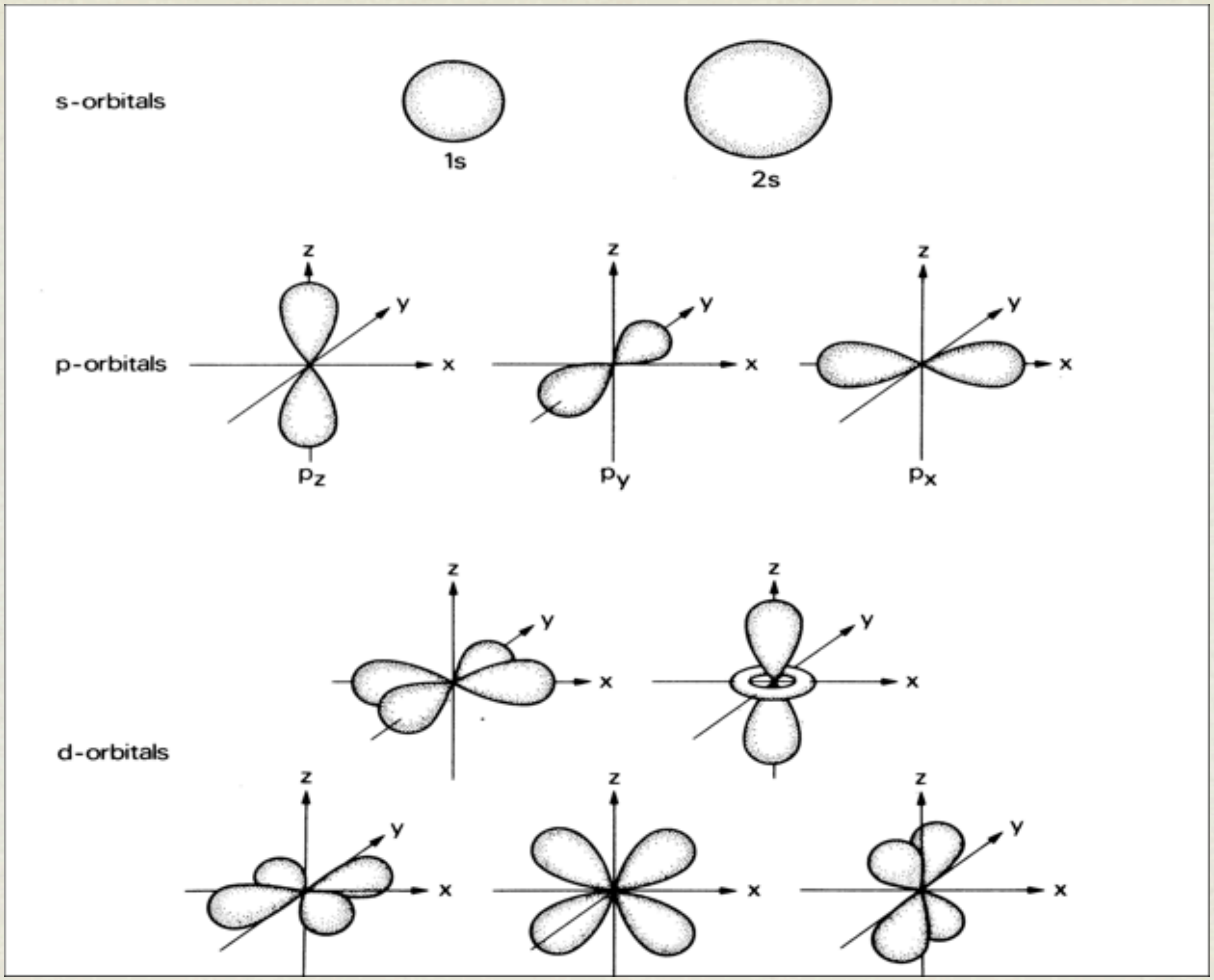
- Each energy level (think: shell) has a number called the PRINCIPAL QUANTUM NUMBER, n
- n can be 1-7 (because there are 7 periods on the periodic table



Secondary Orbital, ℓ

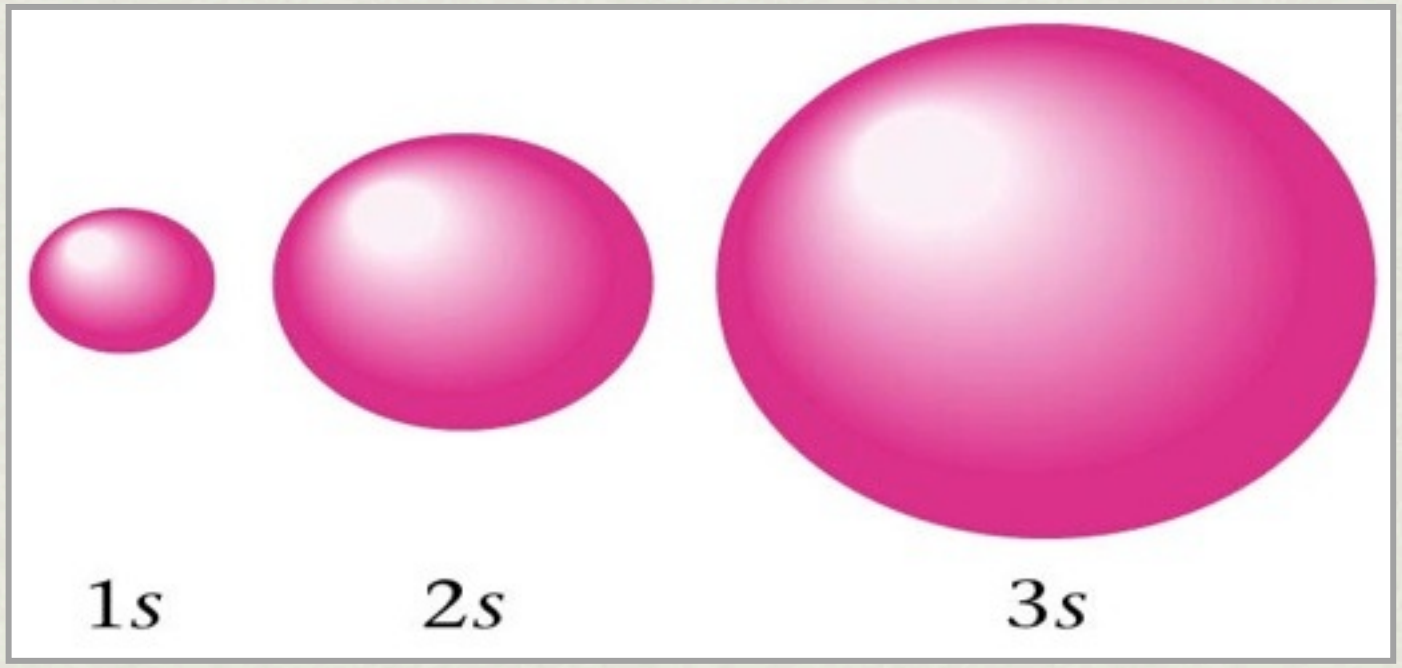
- Each energy level (n) has between 1 and 4 sublevels or subshells, ℓ
- Possible values of ℓ = 0 to n-1
- Each sublevel is associated with a particular shape of probability
- 4 shapes: s, p, d, and f

SHAPES



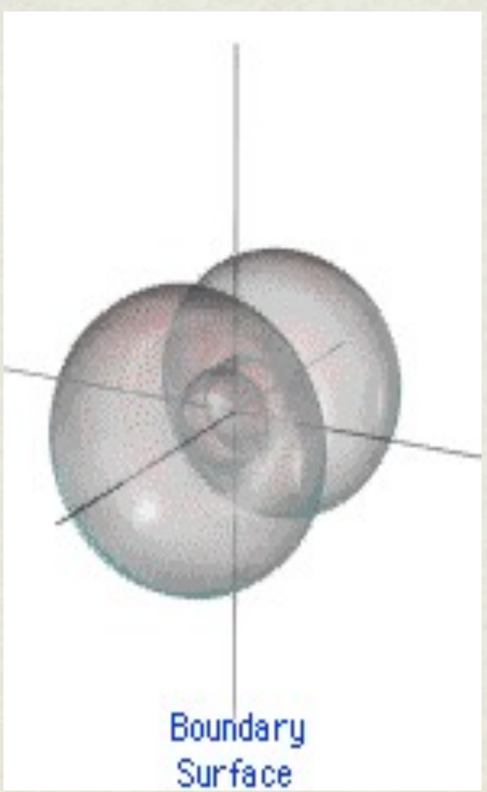
TYPES OF ORBITALS: S

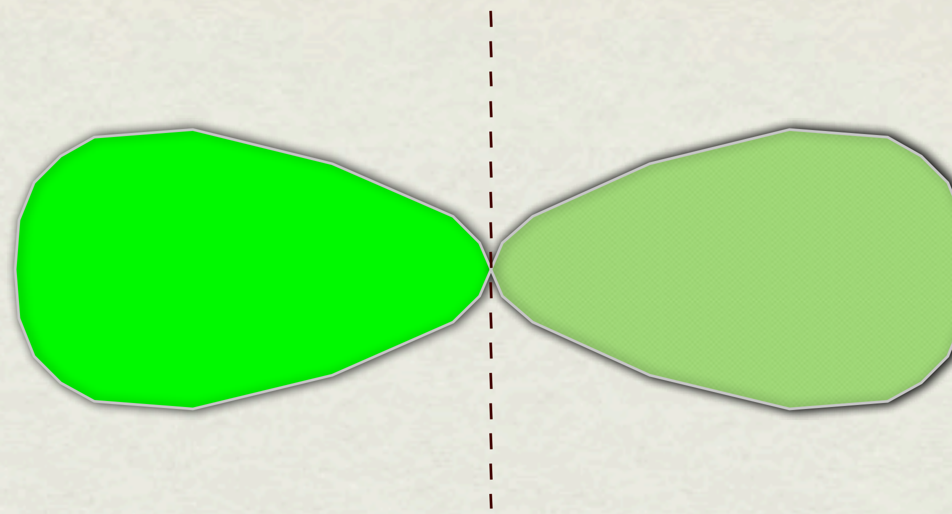
When $\ell = 0$, the orbital is called s



Types of Orbitals (l)

When $\ell = 1$, the orbital is called p

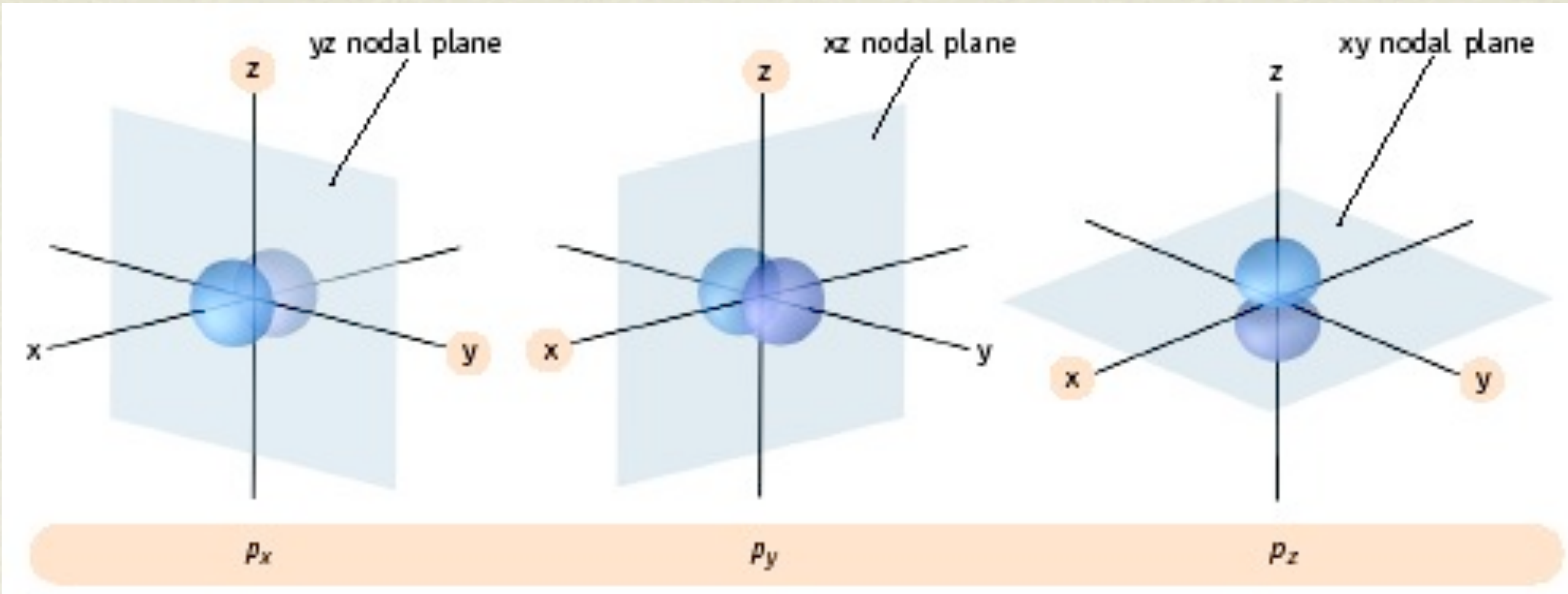




planar node

There is a PLANAR NODE thru the nucleus, which is an area of zero probability of finding an electron

P ORBITALS

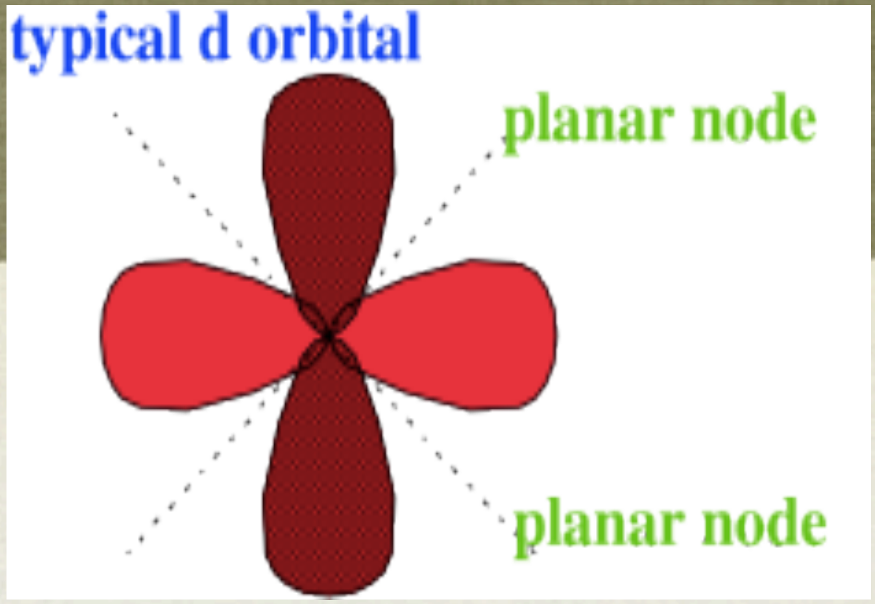


- The p sublevel has 3 orbitals
- The three p orbitals lie 90º apart in space
- They are designated p_x, p_y, p_z for the axis

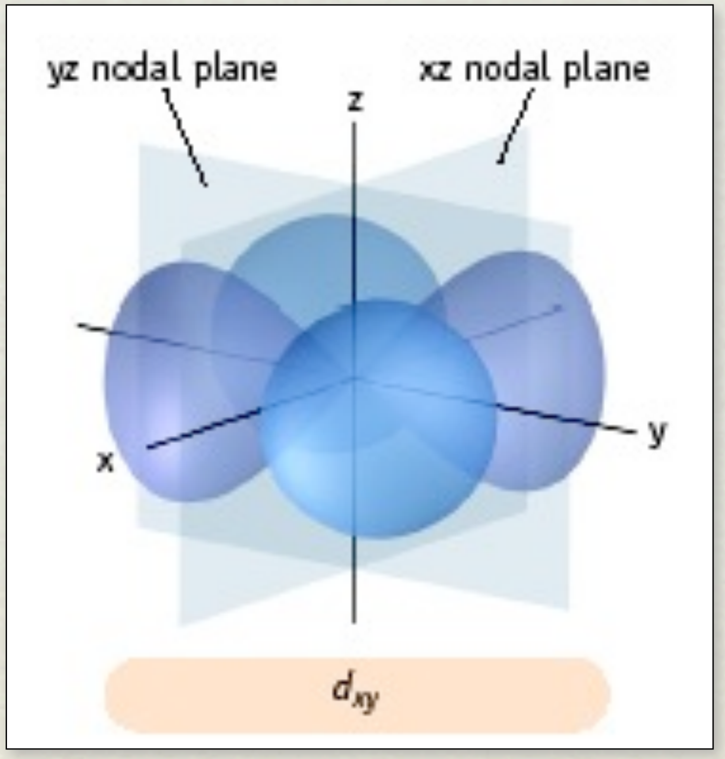
Magnetic (m_l)

- How do we know p orbitals have 3 orbitals?
- Magnetic number represents the orientation

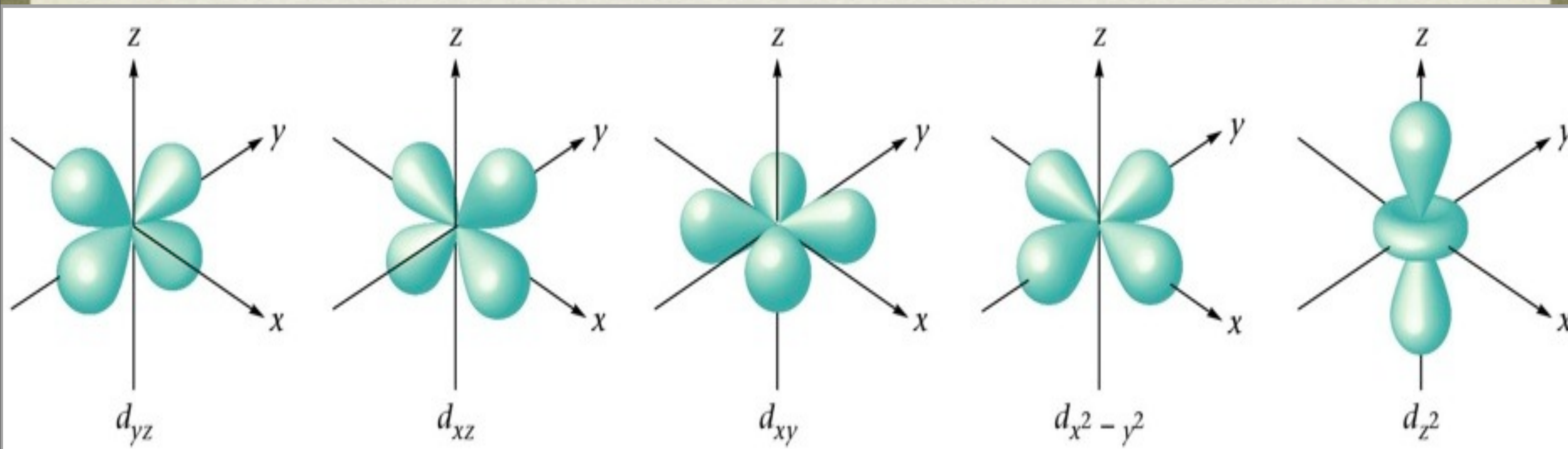
D ORBITALS



When ℓ = 2, the orbital is called d



d orbitals



F ORBITALS

When $l = 3$, the orbital is called f

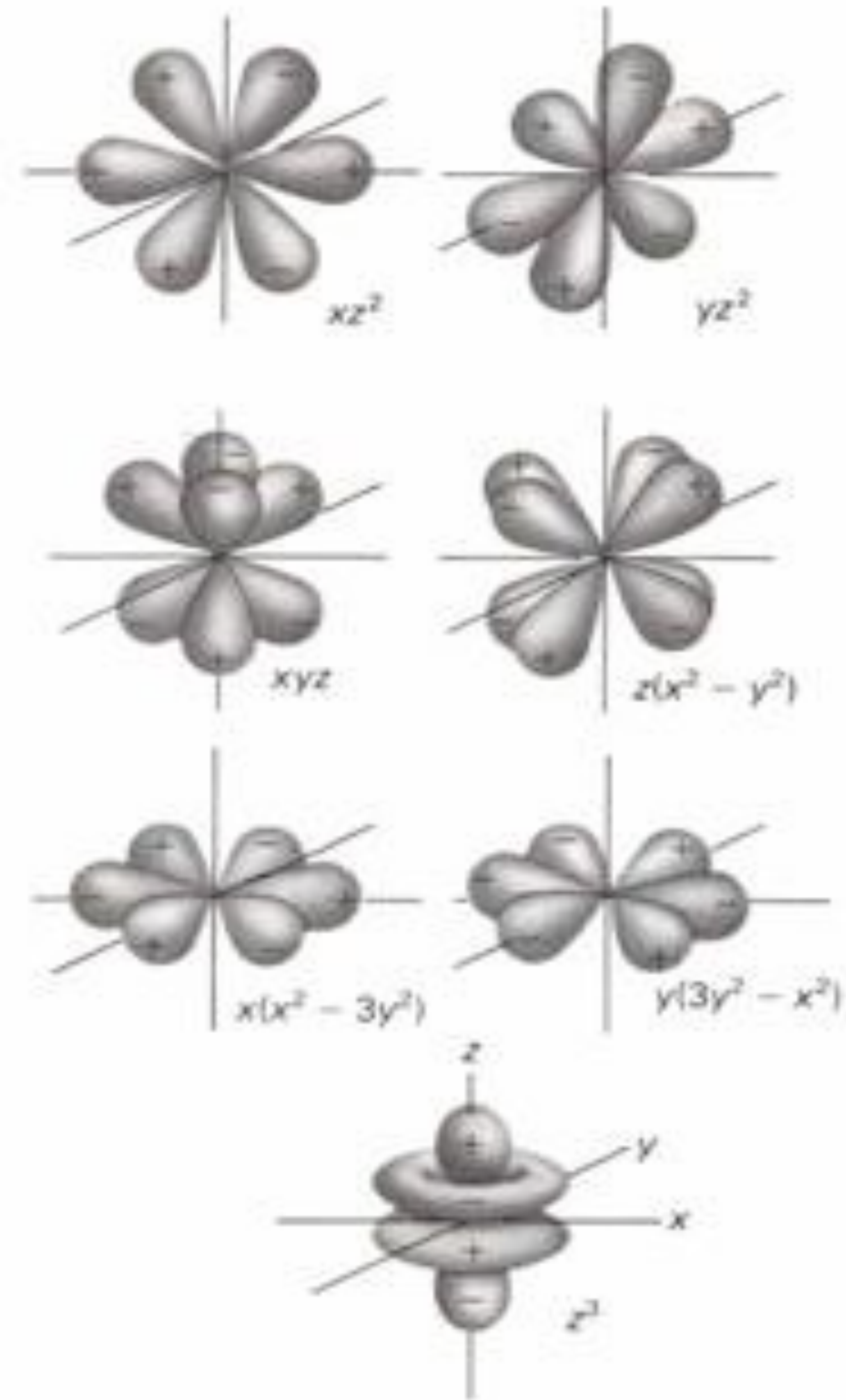


Table 7.1 • Summary of the Quantum Numbers, Their Interrelationships, and the Orbital Information Conveyed

Principal Quantum Number	Angular Momentum Quantum Number	Magnetic Quantum Number	Number and Type of Orbitals in the Subshell
Symbol = n *Values* = 1, 2, 3, ... n = *number of subshells*	*Symbol* = ℓ *Values* = $0 \ldots n-1$	*Symbol* = m_ℓ *Values* = $-\ell \ldots 0 \ldots +\ell$	*Number of orbitals in shell* = n^2 *and number of orbitals in subshell* = $2\ell + 1$
1	0	0	one 1*s* orbital (one orbital of one type in the $n = 1$ shell)
2	0 1	0 +1, 0, −1	one 2*s* orbital three 2*p* orbitals (four orbitals of two types in the $n = 2$ shell)
3	0 1 2	0 +1, 0, −1 +2, +1, 0, −1, −2	one 3*s* orbital three 3*p* orbitals five 3*d* orbitals (nine orbitals of three types in the $n = 3$ shell)
4	0 1 2 3	0 +1, 0, −1 +2, +1, 0, −1, −2 +3, +2, +1, 0, −1, −2, −3	one 4*s* orbital three 4*p* orbitals five 4*d* orbitals seven 4*f* orbitals (16 orbitals of four types in the $n = 4$ shell)

RULES TO FOLLOW

Only 2 electrons per orbital!

The Pauli Exclusion Principle says that no two electrons within an atom (or ion) can have the same four quantum numbers.

If two electrons are in the same energy level, the same sublevel, and the same orbital, they must have opposite spins!

So M_s can be either +1/2 or -1/2 (and can be arbitrarily assigned)

Try it!

- Write a set of quantum numbers for an electron in a **1s** orbital

1s

Try some more!

- Write a set of quantum numbers for an electron in a **3p** orbital

3p

You made it!

- Let’s try writing some quantum numbers!
- Write a set of 4 quantum numbers for an electron in:
- 2s
- 3d
- 4f

Try it!

1. What are the allowed values for l in each of the following cases?

   (a) $n = 5$ (b) $n = 1$

2. What are the allowed values for m_l, for an electron with the following quantum numbers:

   (a) $l = 4$ (b) $l = 0$

3. What are the names, m_l values, and total number of orbitals described by the following quantum numbers?

   (a) $n = 2, l = 0$ (b) $n = 4, l = 3$

4. Determine the n, l, and possible m_l values for an electron in the $2p$ orbital.

5. Which of the following are allowable sets of quantum numbers for an atomic orbital? Explain your answer in each case.

   (a) $n = 4, l = 4, m_l = 0$ (c) $n = 2, l = 0, m_l = 0$

   (b) $n = 3, l = 2, m_l = 1$ (d) $n = 5, l = 3, m_l = -4$

Try it!

- Practice Quantum numbers: p. 184 #3,4,6,7

Summarizing the Four Quantum Numbers for Electrons in Atoms

Quantum Number Name	Symbol	Allowed Values	Property
principal	n	positive integers (1, 2, 3, etc.)	orbital size and energy
orbital-shape	l	integers from 0 to $(n - 1)$	orbital shape
magnetic	m_l	integers from $-l$ to $+l$	orbital orientation
spin	m_s	$+\frac{1}{2}$ or $-\frac{1}{2}$	electron spin direction