Lesson 1

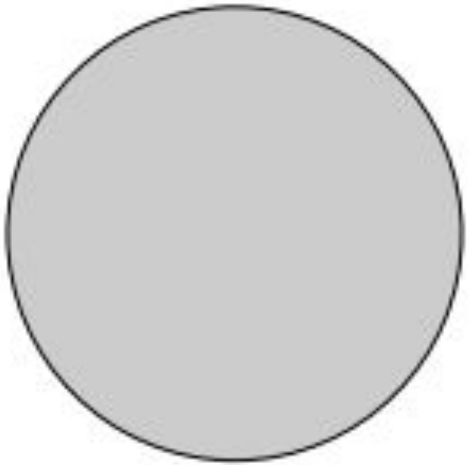
Atomic Structure

Learning Goals

- ❑ Review early atomic theory and analyze contributions made by atomic theorists as they apply to modern theory of the atom.
- ❑ Identify properties of electron of atom based on a set of four quantum numbers.

Dalton’s Atom (1808)

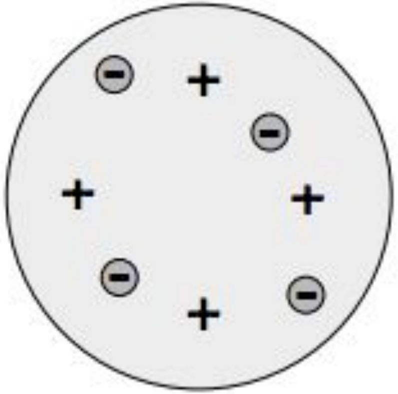
- Atoms are indivisible spheres.



John Dalton
(1766-1844)

Thomson’s Plum Pudding Model (1904)

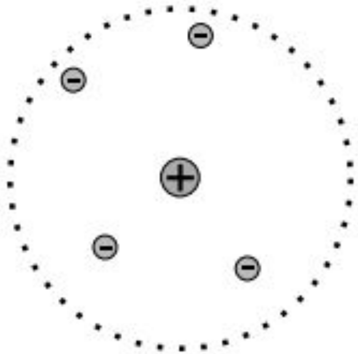
- Based on his discovery of electron in 1897.
- The atom is mostly a spheres of positive charge.
- Tiny, negatively charged “electrons” circulate in this sphere.
- The positive and negative charges balance - atom is neutral.



J. J. Thomson
(1856-1940)

Rutherford’s Nuclear Model (1911)

- Based on the results of the gold-foil experiment.
- Most of the atom is empty space.
- Atom’s mass is concentrated in a tiny positive nucleus.
- Electrons are located in the space surrounding the nucleus.



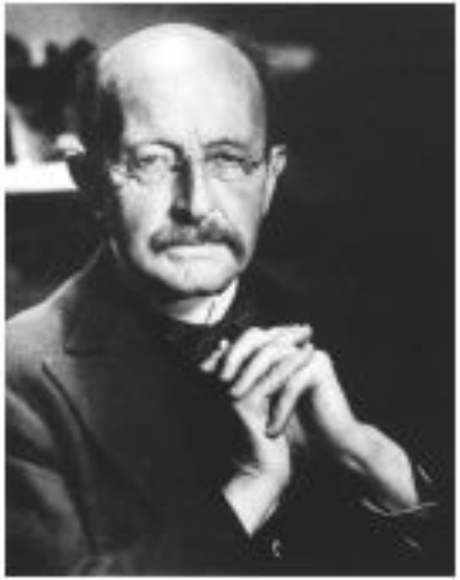
Ernest Rutherford
(1871-1937)

Thomas Young
(1773-1829)

- Light shows wave interference.
- Light behaves like a wave.

- Light is an electromagnetic wave.

James Maxwell
(1831-1879)



Max Planck
(1858-1947)

- Light from hot objects is emitted as discrete packets of energy (quanta).
- The energy of a quantum depends on the frequency of the light.

$$E = hf$$

- Light is absorbed as quanta (explains the photoelectric effect).
- The energy of a quantum depends on the frequency of the light.

$$E = hf$$

PHOTONS

Albert Einstein
(1879-1955)

Bohr's Quantum Atomic Model (1913)

- Electrons orbit the nucleus only at specific allowable energy levels (stationary states).
- While in an stationary state, electrons do not emit energy.
  - Electrons cannot spiral into the nucleus
- An electron can change to a higher energy level by absorbing a photon with energy exactly equal to the difference between the energy levels.
  - Explains the absorption spectra (dark-line spectral) of elements
- An electron can change to a lower energy level by emitting a photon with energy exactly equal to the difference between the energy levels.
  - Explains the emission spectra (bright-line spectra) of elements

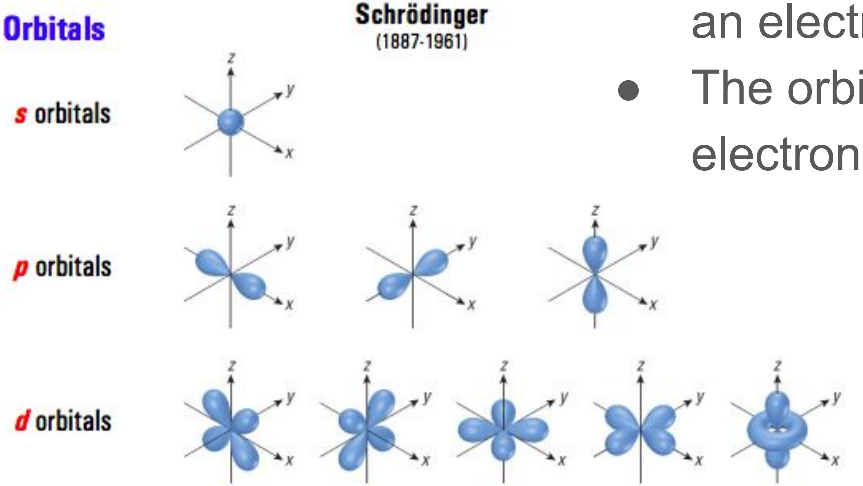
De Broglie's Matter Waves (1924)

- Hypothesizes that all particles (including electrons) have wave properties.
- Later supported by observed wave interferences patterns produced by electrons.

Louis
de Broglie
(1892-1987)

Schrodinger's Wave Mechanics (1927)

Erwin
Schrödinger
(1887-1961)



- Applies wave equations to the electrons in an atom.
- Solution describe probability distribution for the electrons called orbitals.
- An orbital is a three-dimensional region in which an electron is most likely found.
- The orbital does not give the exact positive of the electron or the motion of the electron.

Heisenberg's Uncertainty Principle (1927)

- The exact position and motion of an electron (or any particle) cannot be known.

Werner
Heisenberg
(1901-1976)

The Quantum Numbers

These numbers are used to describe the probable location of an electron.

1. **The Principal Quantum Number (*n*)**
   The value for the energy level or **orbital** of an electron ($n = 1, 2, 3$ *to* ∞).
2. **The Secondary Quantum Number (*l*)**
   Divides the energy levels into **sublevels** or **orbital types** (smaller spaces or shapes) called *s (0), p (1), d (2), f (3)*.
   $l = 0$ **to** $(n-1)$.
3. **The Magnetic Quantum Number (m_l)**
   Describes the **direction of orientation** (pointing) for *p, d and f* orbital types (i.e. p_x, p_y, p_z). $m_l = -l$ *to* $+l$
4. **The Spin Quantum Number (m_s)**
   An electron can spin in only 2 directions
   ($m_s = +½$ **or** $-½$)

Quantum Numbers Questions

1) An electron has the following set of quantum numbers:
   $n = 3$, $l = 1$, $m_l = 1$, $m_s = -\frac{1}{2}$.
   In which orbital is this electron found?
   a. 3*s* b. 3*p* c. 3*d* d. 3*f* e. 4*p*

2) Which set of quantum numbers is not possible?
   a. $n = 3$, $l = 0$, $m_l = 0$, $m_s = -\frac{1}{2}$
   b. $n = 5$, $l = 3$, $m_l = 2$, $m_s = \frac{1}{2}$
   c. $n = 4$, $l = 3$, $m_l = -1$, $m_s = -\frac{1}{2}$
   d. $n = 5$, $l = 3$, $m_l = -3$, $m_s = \frac{1}{2}$
   e. $n = 4$, $l = 4$, $m_l = 2$, $m_s = -\frac{1}{2}$

3) Which set of quantum numbers is not possible?
   a. $n = 5$, $l = 3$, $m_l = 0$, $m_s = -\frac{1}{2}$
   b. $n = 1$, $l = 0$, $m_l = 0$, $m_s = -\frac{1}{2}$
   c. $n = 3$, $l = 2$, $m_l = 1$, $m_s = \frac{1}{2}$
   d. $n = 4$, $l = 3$, $m_l = -3$, $m_s = 0$
   e. $n = 5$, $l = 2$, $m_l = 0$, $m_s = \frac{1}{2}$

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Quantum Numbers Questions

4) What are the allowed values of m_l for an electron with each orbital-shape quantum number.

   a) $l = 3$ b) $l = 1$

5a) What are the possible values of m_l if $n = 4$ and $l = 2$?

b) What kind of orbital is described by these quantum numbers?

c) How many orbitals can be described by these quantum numbers?

Success Criteria

- ❑ I can explain how experimental observations and inferences by Ernest Rutherford and Niels Bohr contributed to the development of the planetary model of the atom.

- ❑ I can describe the quantum mechanical model of the atom in terms of shells (energy levels), subshells (sublevels), orbitals, and electron spin.