Chapter 4
Electron Configurations

Waves

- Today scientists recognize light has properties of waves and particles
- Waves: light is electromagnetic radiation and travels in electromagnetic waves.

4 Characteristics of a wave:

- **1) amplitude - height of the wave. For light it is the brightness**
- **2) Wavelength (λ)– distance from crest to crest.**
  - **For light – defines the type of light**
  - **Visible light range – 400-750nm**

Properties continued

- 3) Frequency (ν)– measures how fast the wave oscillates up and down.
  - It is measured in number per second.
  - Hertz = 1 cycle per second
  - Visible light = 4×10^{14} cycles per second to 7×10^{14} cycles per second
- 4) speed – 3.00×10^{8} m/s (MEMORIZE)

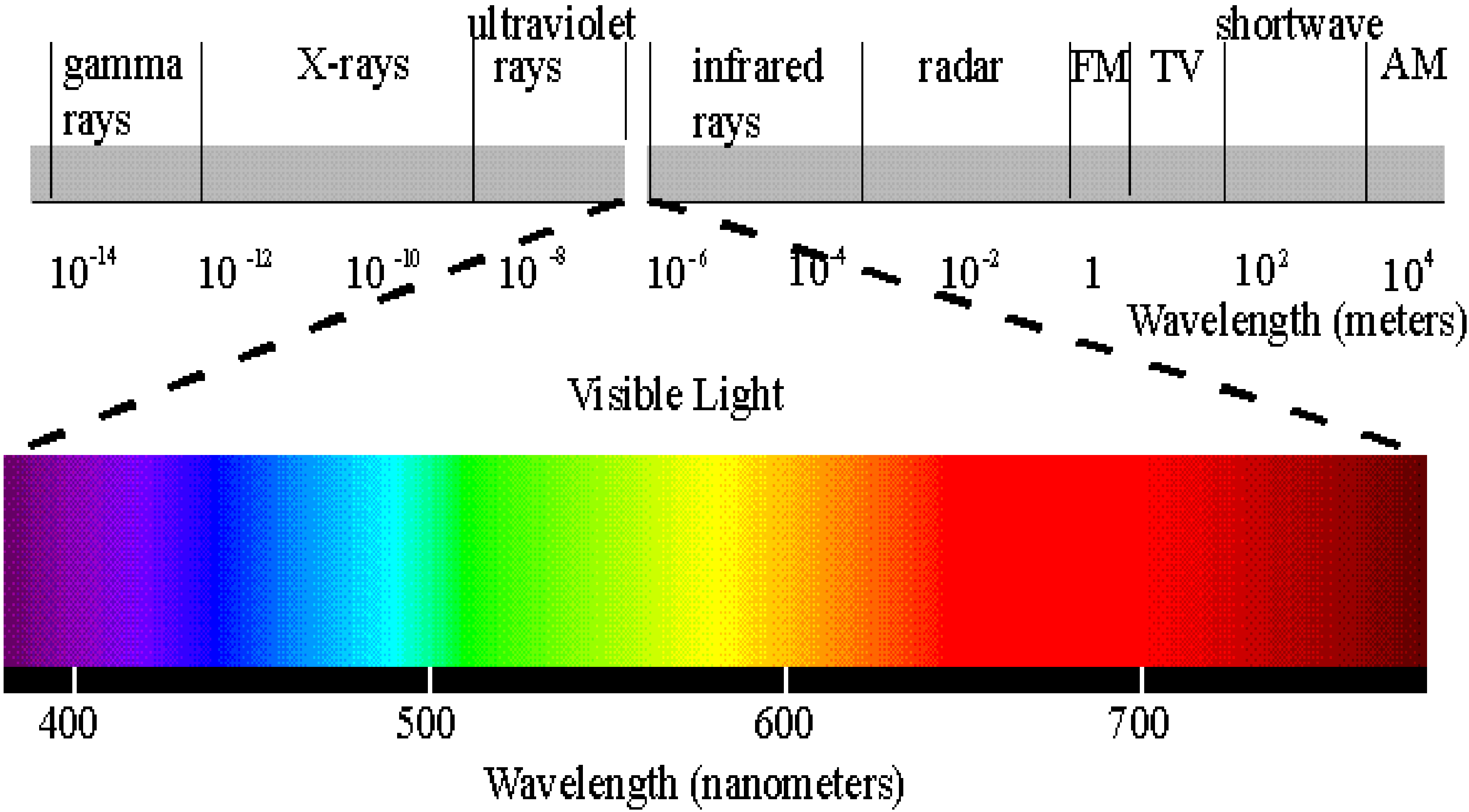
Shedding more light

- Short wavelength, high frequency
- Long wavelength, low frequency
- Visible Spectrum

ROY	G	BIV
Longer wavelength		shorter wavelength

Electromagnetic spectrum (meters)

- 10^{-11} gamma
- 10^{-9} x-rays
- 10^{-8} UV
- 10^{-7} visible light
- 10^{-6} infrared
- 10^{-2} microwave
- 1 TV



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Visual Stimulus

Wavelength and frequency

- **Wavelength and frequency are inversely related!!**

  $$\lambda = c/\nu$$

- **Where λ is the wavelength, c is the speed of light and ν is the frequency**
- **Speed of light = Constant =**

  3.00×10^8 m/sec

Example

- Example: An infrared light has a wavelength of 2.42×10^{-6}m. Calculate the frequency of this light.
- $\nu = c/\lambda$
- $\nu = \dfrac{3.0 \times 10^{8}\text{m/sec}}{2.42 \times 10^{-6}\text{m}} =$
- $= 1.2 \times 10^{14}$ waves/sec

Wavelength and frequency

- ****Remember λ and ν are inverse. Therefore short wavelength = high frequency!!

Atom History

- Atoms solid balls
- P^+, n^0, and e^- …nuclear atom
- Solar System atom
- Bohr atom…H only
- Quantum model…explains why elements when heated give off unique wavelengths of light (flame test)

1900 Quantum Theory

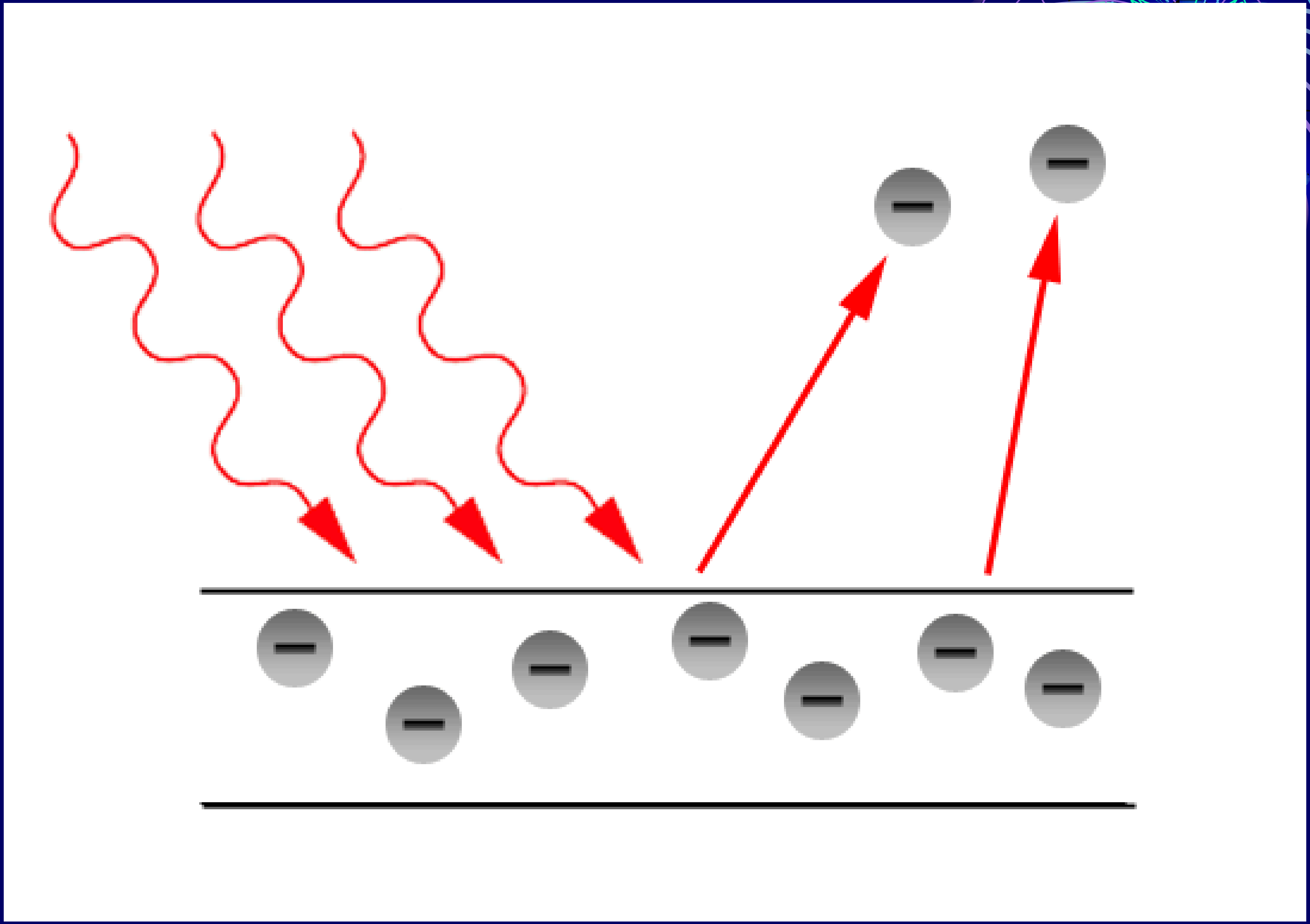
Max Planck proposed the idea

- The amount of energy an object absorbs/emits occurs only in fixed amounts called quanta (quantum)
- Quanta – discrete amount of energy that can be gained or lost by an atom/electron

1905 Einstein's theory

- Einstein proposed that light (because it is energy) consists of quanta of energy called PHOTONS

- Photon = discrete bit of light energy



Photoelectric effect –

- Electrons are ejected from the surface of a metal when light shines on the metal.

- The wavelength and frequency determines the amount of energy.

- The higher the frequency, the more energy per photon.

Energy equation

- Amount of energy of a photon described as

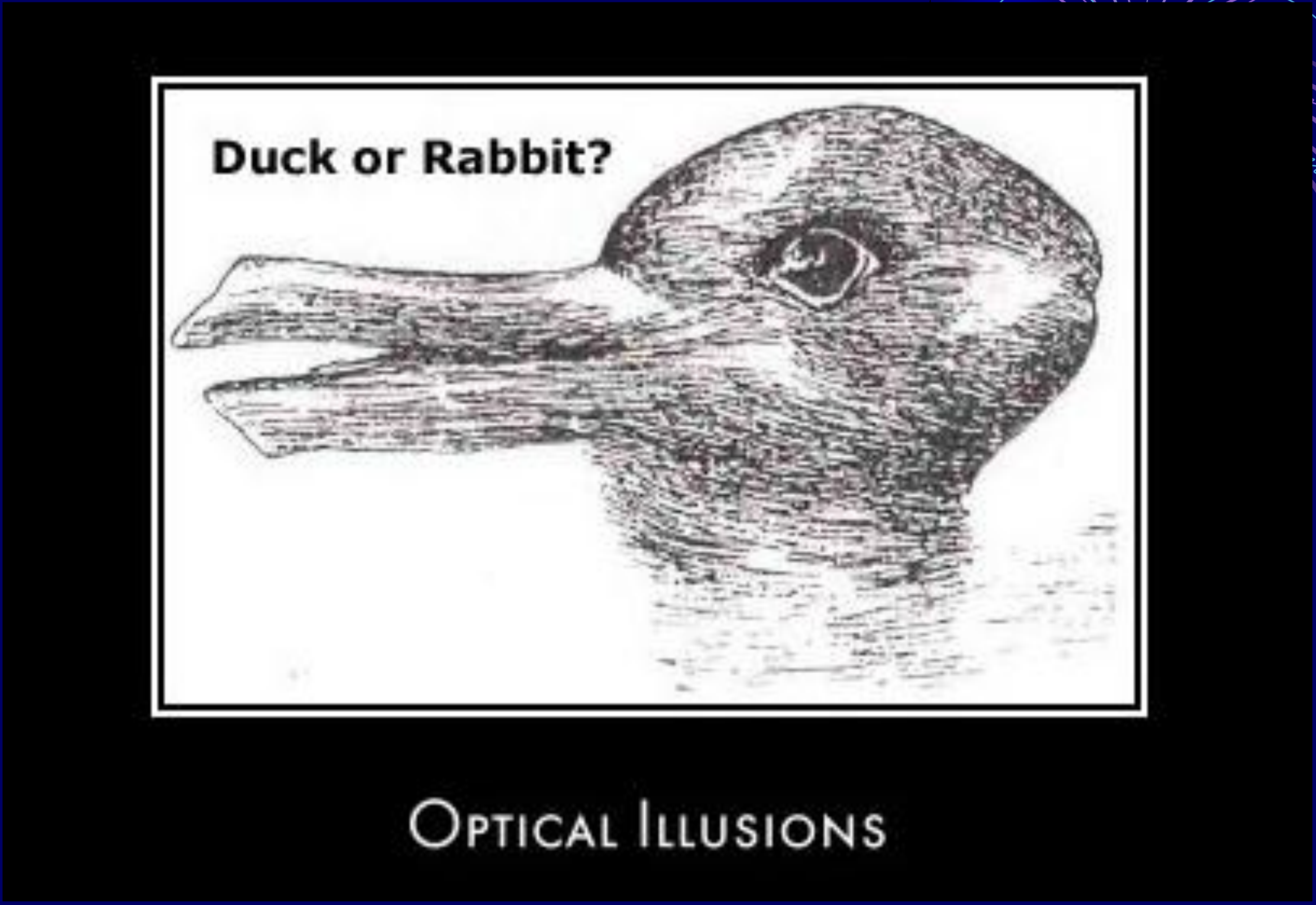
$$E = h\nu$$

E = energy

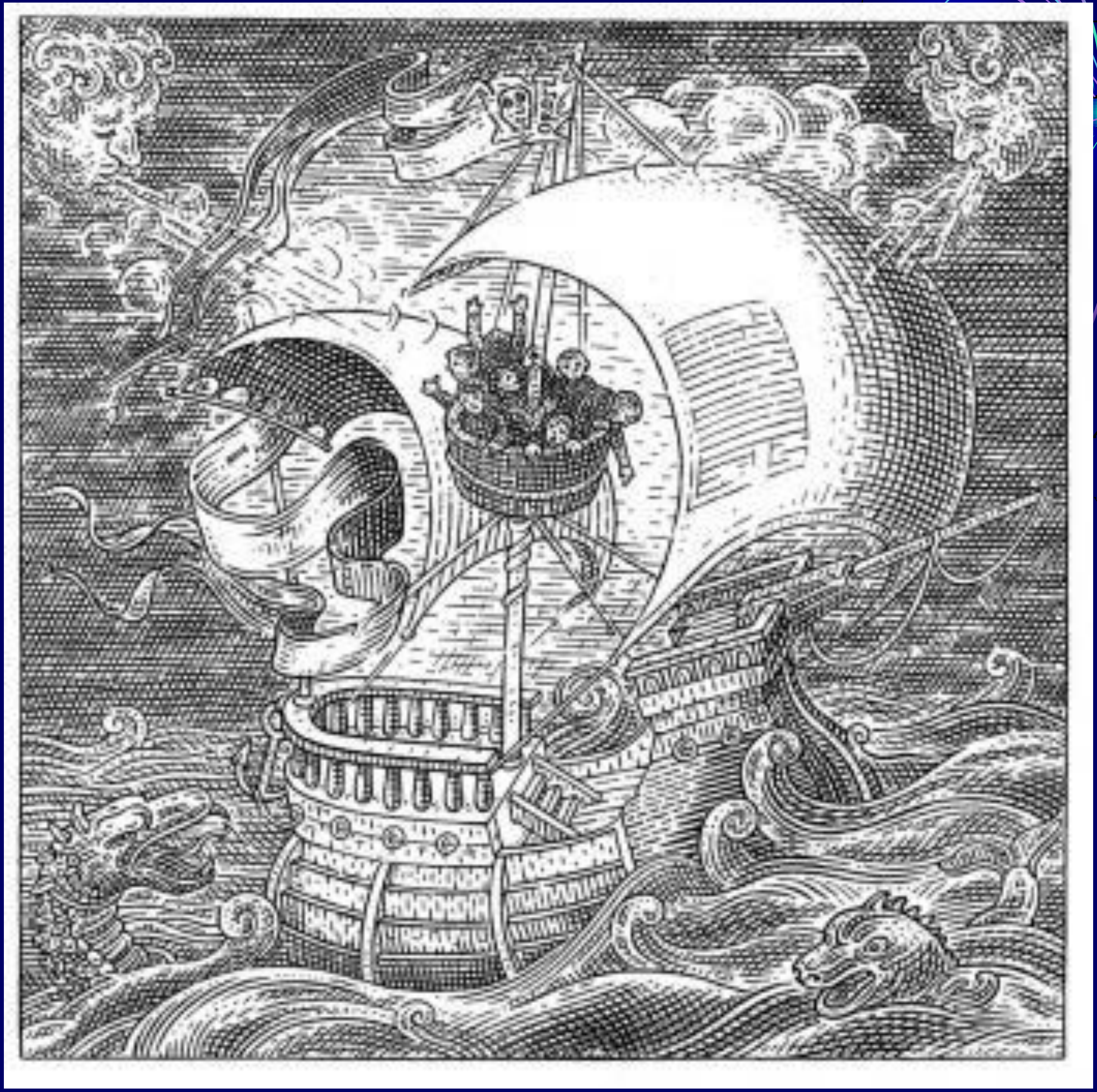
ν = frequency

h = Planck's constant = 6.626×10^{-34} J s

- Joule = SI unit for energy


Duck or Rabbit?
OPTICAL ILLUSIONS


YOUNG WOMAN OR OLD LADY?



Dual nature of radiant energy

- Photons act BOTH like particles and waves.

Studying atoms using light

- All elements emit light when they are energized
- Bright Line Spectra: A spectrum that contains only certain colors, or wavelengths

How are electrons arranged in atoms?

- **Explanation: Bohr atom: 1911**
  - **postulated that the electrons orbit in rings called *energy levels***
- **energy levels are labeled by a quantum number, n.**
- **lowest energy level n=1**
  - **Called ground state**

How are electrons arranged in atoms?

- electron absorbs energy, it jumps to a higher level (known as the excited state) n = 2 or 3 or 4
- Bohr model of an atom
- Only worked for Hydrogen

1924 – Louis DeBroglie

- If waves of light can act as a particle, then particles of matter should act like a wave. Found to be true.

DeBroglie

- Matter waves = wavelike behavior of particles.
- Wave nature is inversely related to mass so we don’t notice wave nature of large objects.
- However, electrons have a small mass and the wave characteristic is more noticable

Schroedinger’s wave equation

- predicted probability of finding an electron in the electron cloud around nucleus.
- Gave us four numbers to describe the “position”.

Heisenberg's Uncertainty Principle

- The position and momentum of a moving object cannot simultaneously be measured and known exactly.
- Cannot know where it is and where its going at the same time.

Quantum mechanical model of an atom –

- Treats the electrons as a wave that has quantized its energy

- Describes the probability that electrons will be found in certain locations around the nucleus.

“Locating” an electron…

- What is your address?

- Four parts of your address…
  - State
  - City
  - Road
  - House number

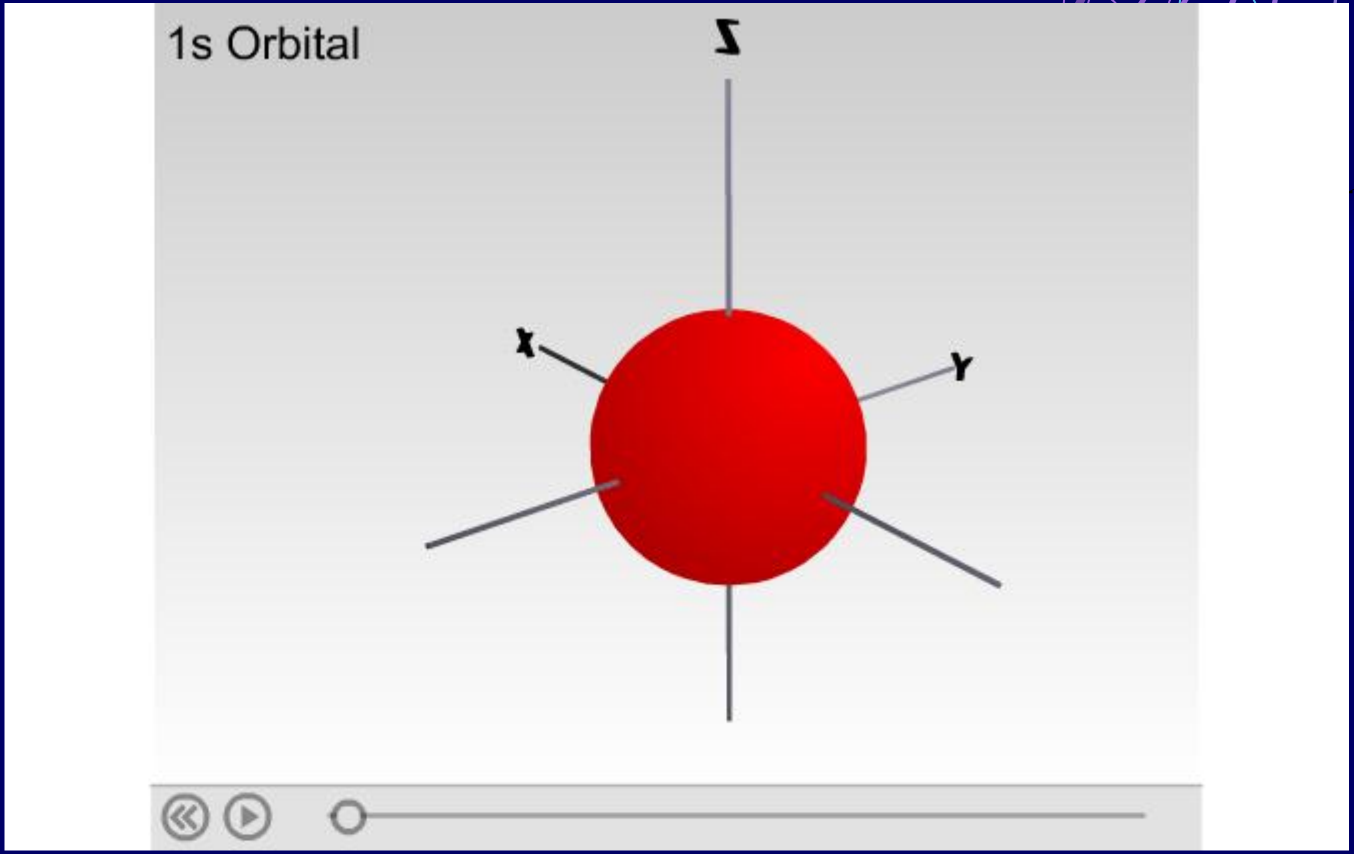
Energy Level (State)

- n=1 … n=any whole number
- Describes which “ring”
- Indicates
  - amount of energy
  - size of region
  - distance from the nucleus
- Higher the number the higher all of the above will be

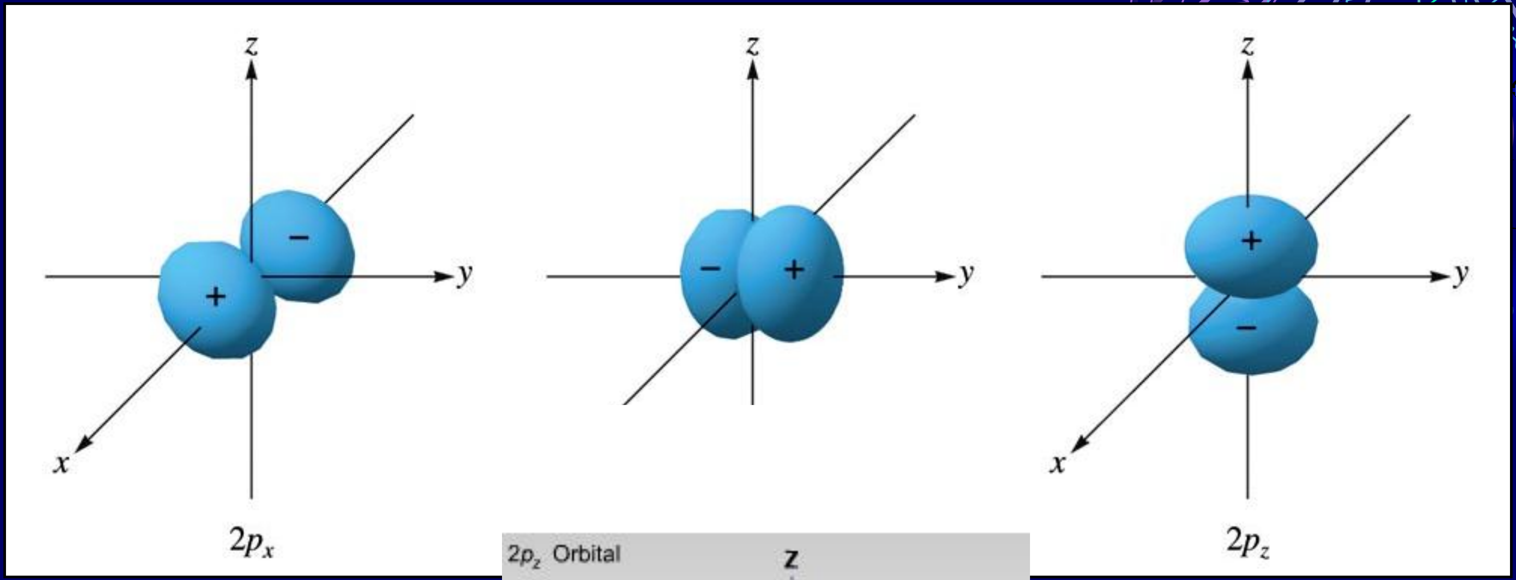
Sublevel (City/Town)

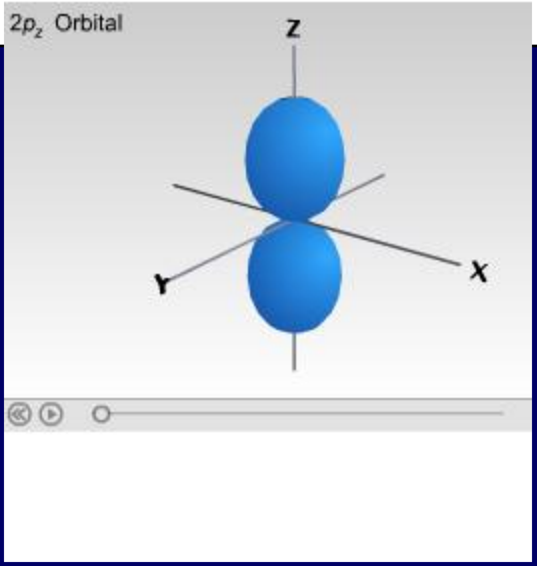
- Division of energy level
  - Number of sublevels = n
  - n=1…1 sublevel
  - n=2…2 sublevels etc…
- Sublevels have characteristic shapes
- Four different kinds of sublevels
  - s, p, d, f (each is a different shape)

S sublevel

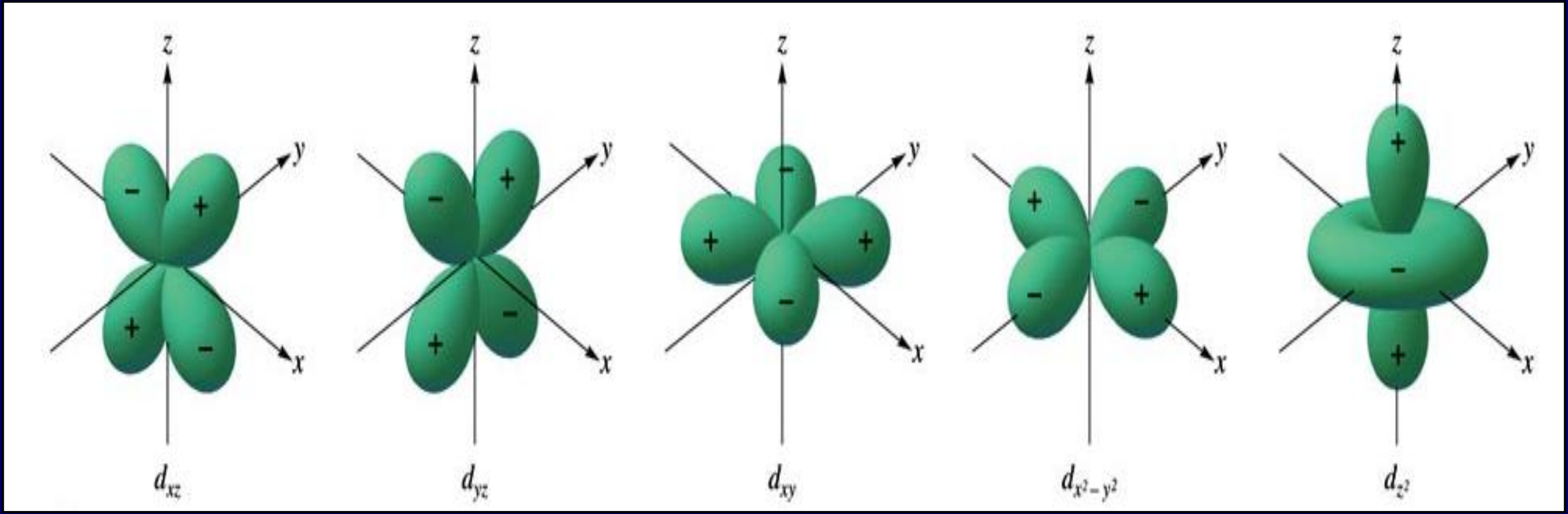


p sublevel

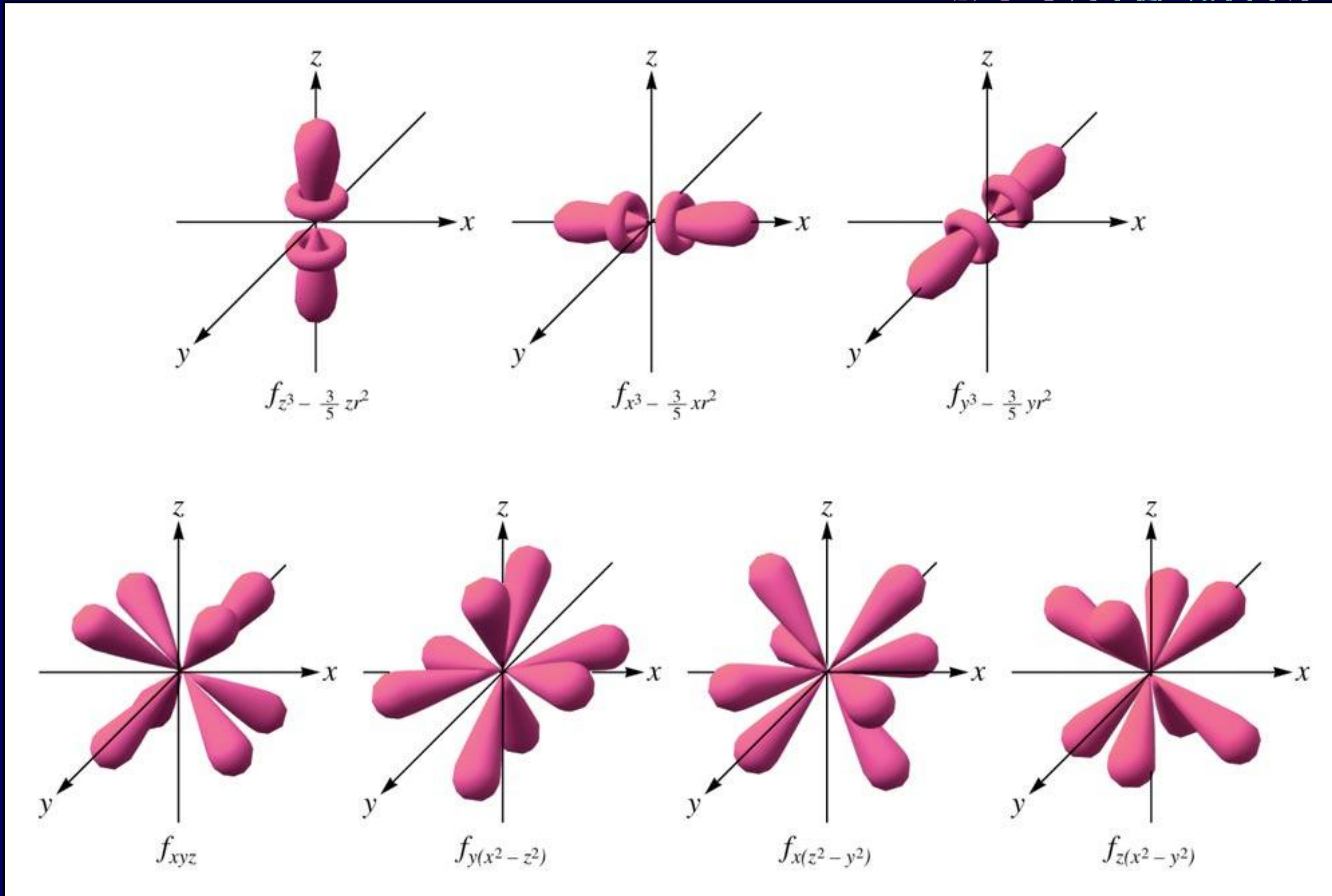




d sublevel



f sublevel



z
x
y
$f_{z^3-\frac{3}{5}zr^2}$
$f_{x^3-\frac{3}{5}xr^2}$
$f_{y^3-\frac{3}{5}yr^2}$
f_{xyz}
$f_{y(x^2-z^2)}$
$f_{x(z^2-y^2)}$
$f_{z(x^2-y^2)}$

Orbital (Street/Road)

- Each sublevel has a certain number of orbital arrangements three dimensionally around the nucleus
  - s = 1 sphere
  - p = 3 (along the x,y or z axis)
  - d = 5
  - f = 7

Spin direction (house #)

- Each electron in an orbital will have a spin
  - 2 options clockwise vs. counter clockwise.
- Pauli Exclusion Principle – each orbital in an atom can hold a maximum of 2 electrons and their electrons must have opposite spin.

Let's see if we get it…

1. How many orbitals are in the 3p sublevel?
2. How many sublevels are in energy level 2?
3. What are the sublevels in energy level 4?
4. How many orbitals, total, exist in all of energy level 3?

Electron configuration rules

- Aufbau Principle - electrons are added one at a time to the lowest energy position available
- Hund's Rule(s)
  - electrons occupy equal energy orbitals so that the maximum number or unpaired electrons result.
  - Occupy singly before pairing

Orbital Diagrams

- Show electrons as ↑ or ↓
- Follow all rules
- Difference between paired and unpaired electrons
  - Paired = 2 electrons in the same orbital
  - Unpaired = 1 electron in the orbital

Electron Configuration

- Shorter way to show electron "locations"
- Hints:
  - Coefficients – energy level
  - Sublevels – s,p,d,f
  - Superscript – number of electrons (remember limits of each sublevel)

Nobel Gas Shortcut

- Find the closest, lower number nobel gas
- Use symbol for nobel gas in [ ]
- Finish rest of electrons as before
  - Ex. Hg (80 electrons)
    - $[Xe]6s^2 4f^{14} 5d^{10}$