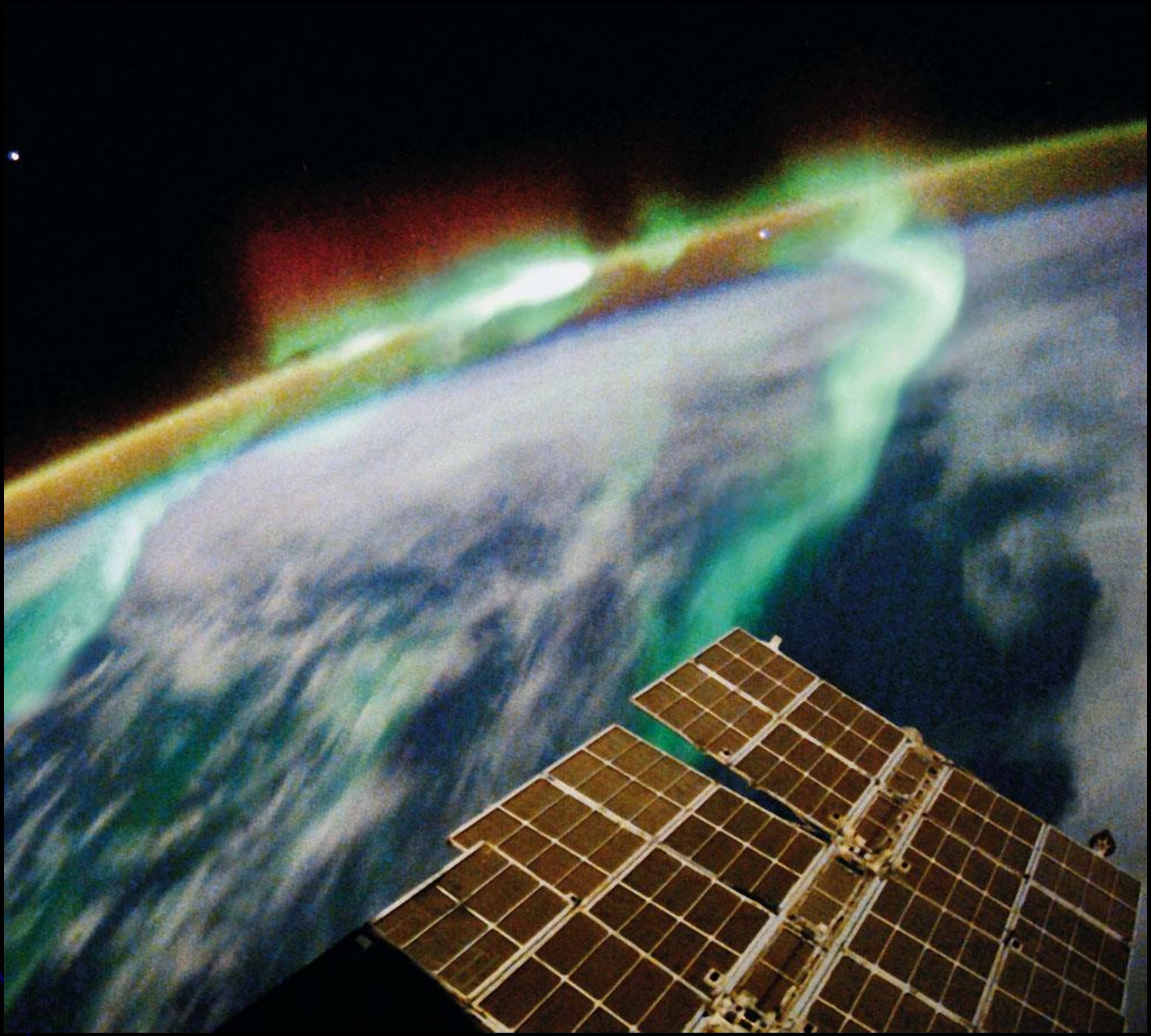
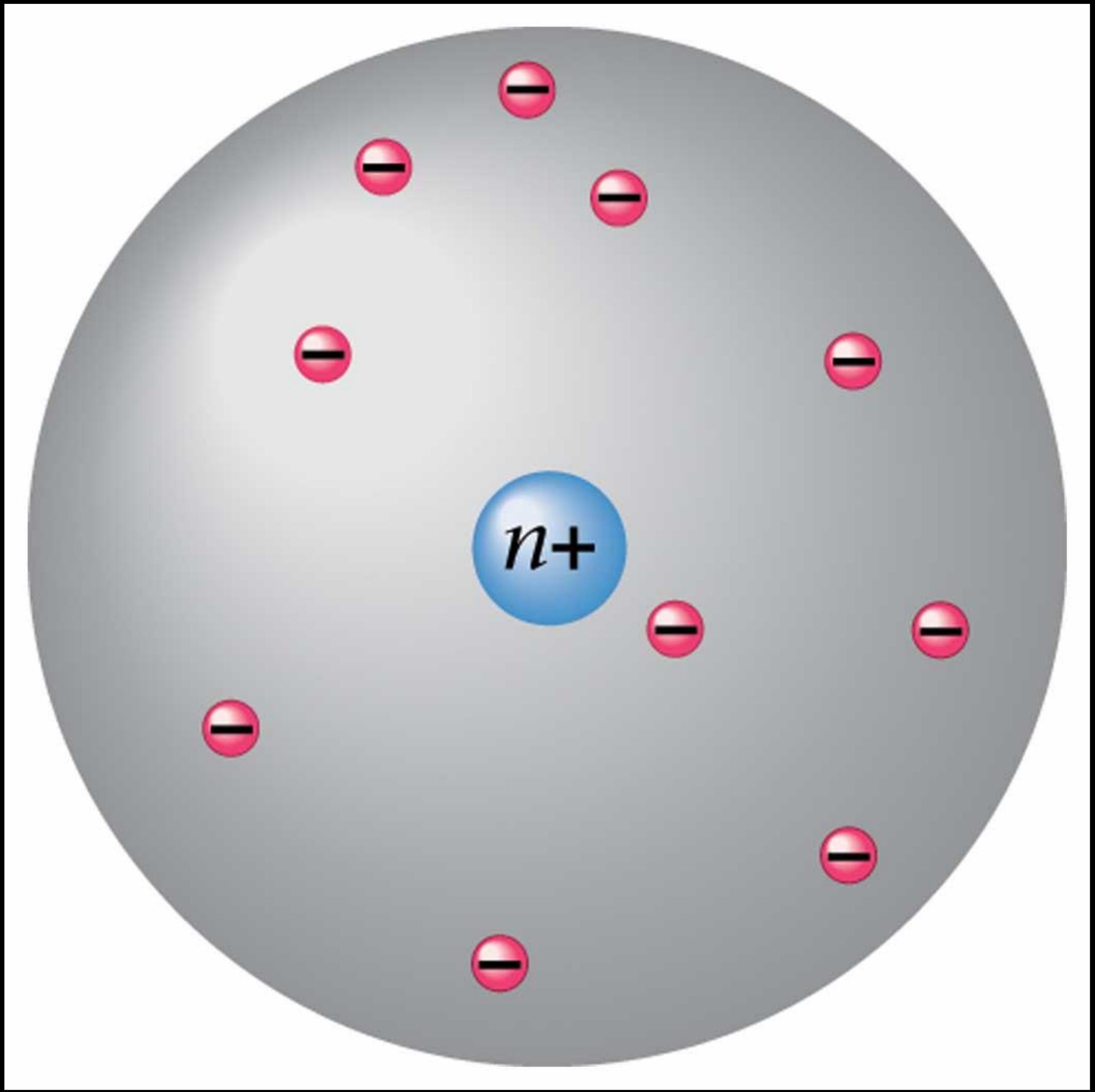
Modern Atomic Theory

Chapter 11



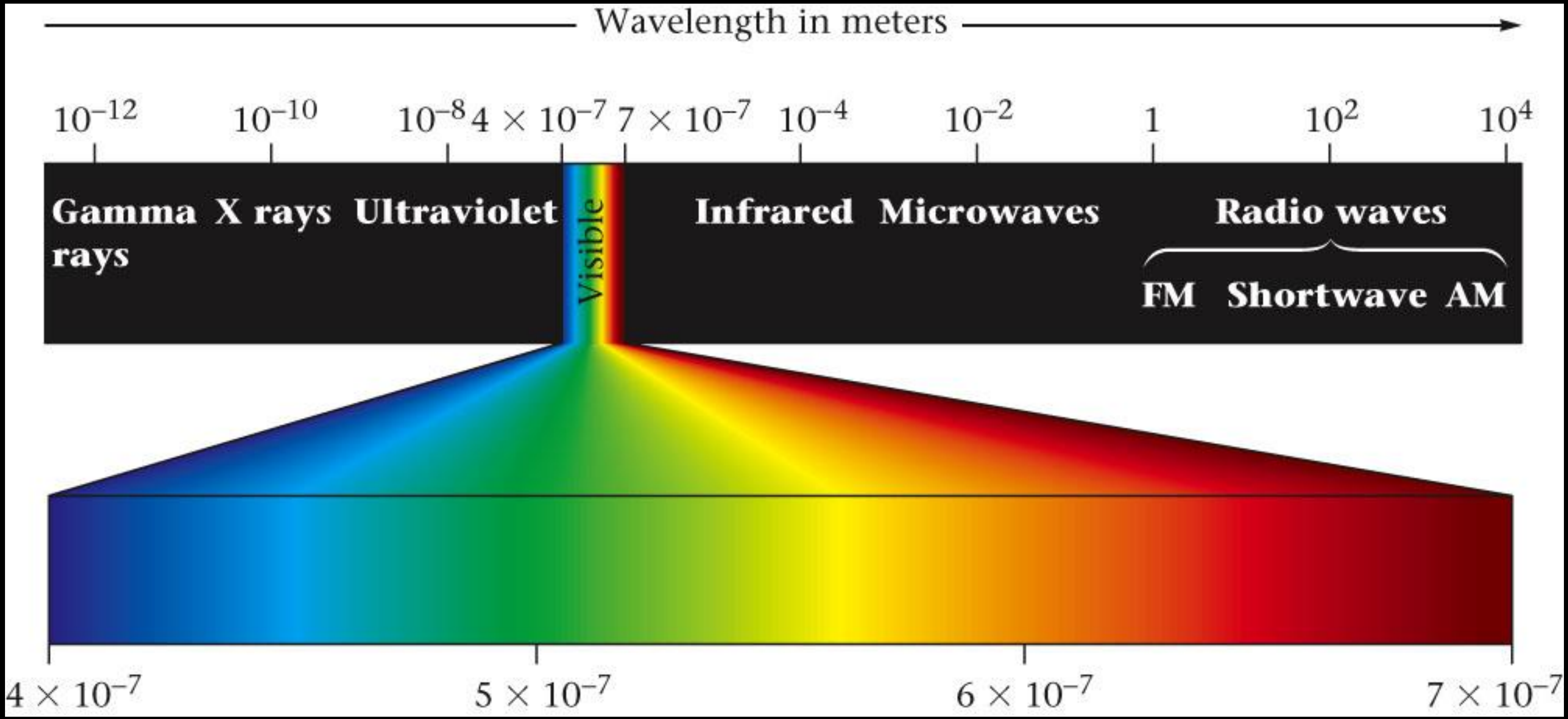
11.1 Rutherford's Atom



11.2 Electromagnetic Radiation

Many types exist all around us.

- **ultraviolet**
- **sunlight (visible light)**
- **x-rays**
- **microwaves**
- **radio waves**
- **infrared**
- **gamma rays**



The different wavelengths of electromagnetic radiation.

What is a wave?

All ER moves at 3.0×10^8 m/s in a vacuum. This is called the **"speed of light."**

Characteristics:

1. **Wavelength: λ = lambda**

   This is the distance between two consecutive peaks or troughs.

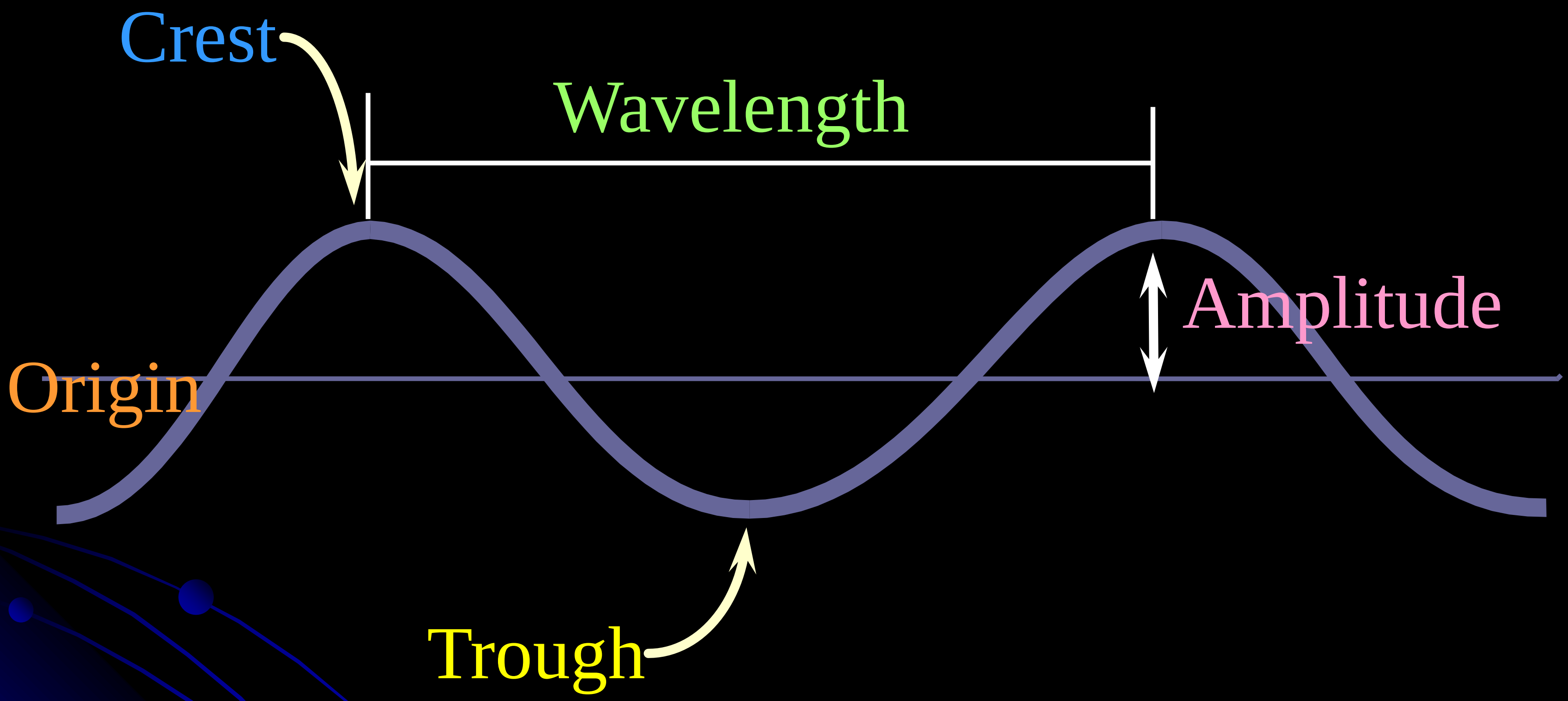
2. **Frequency: ν = nu**

   This is how many waves pass a point per second.

3. **Speed: $c = 3.0 \times 10^8$ m/s (E. R.)**

   This is how fast the wave moves through space.

Parts of a wave



Frequency and wavelength are inversely related.

That means as one increases the other decreases.

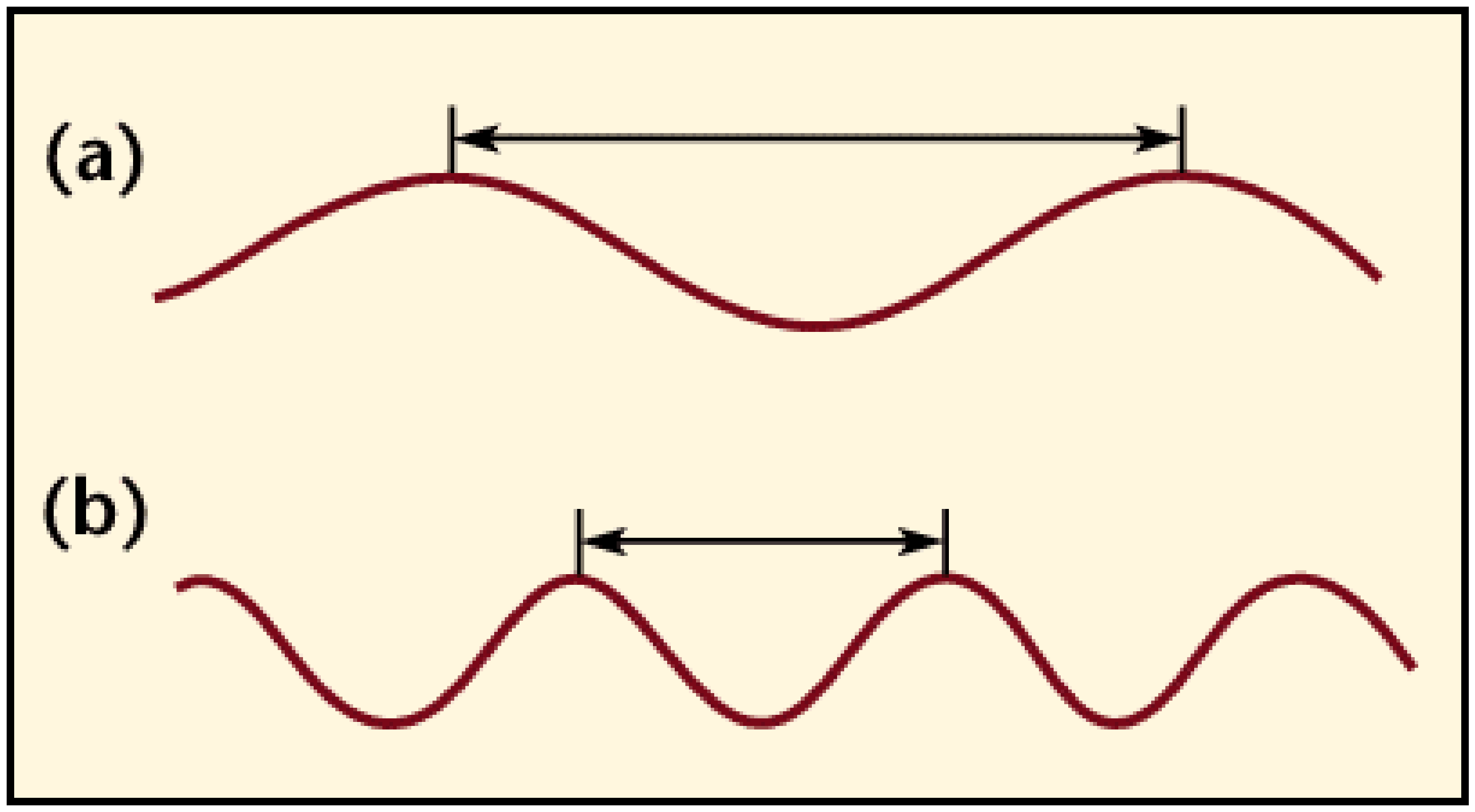
-The longer the wavelength the lower the frequency.

-The shorter the wavelength the higher the frequency.

$$c = \lambda \times \nu$$

Solve for wavelength or frequency.

$$\lambda = \frac{c}{\nu} \quad \text{or} \quad \nu = \frac{c}{\lambda}$$



a. **Longer** wavelength

b. Shorter wavelength

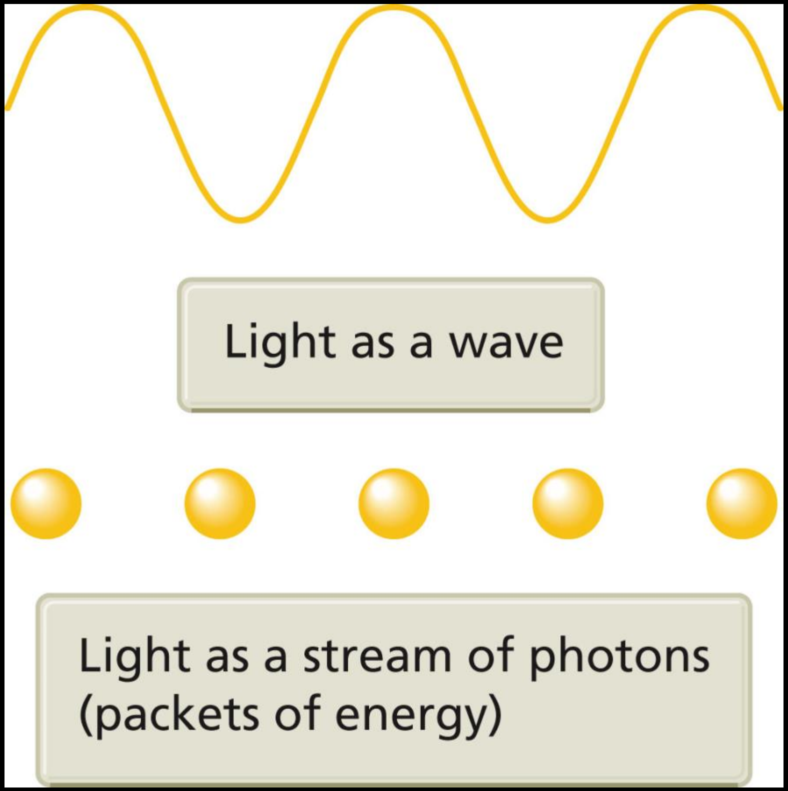
-These waves carry energy.

The *lower* the frequency
the *lower* the energy.

The *higher* the frequency
the *higher* the energy.

- We typically think of light as being transmitted in waves, but it is also thought that light is made up of *packets of energy* called **photons**. (Like little packages.)

- Light seems to behave both **waves** and as **particles**.



11.3 Emission of Energy by Atoms
11.4 The Energy Levels of Hydrogen

-When an atom absorbs energy (electric current, heat, etc.) the atom is in an **excited state**.

-It is unable to maintain this excited state.

-The atom gives off the extra energy in the form of photons.

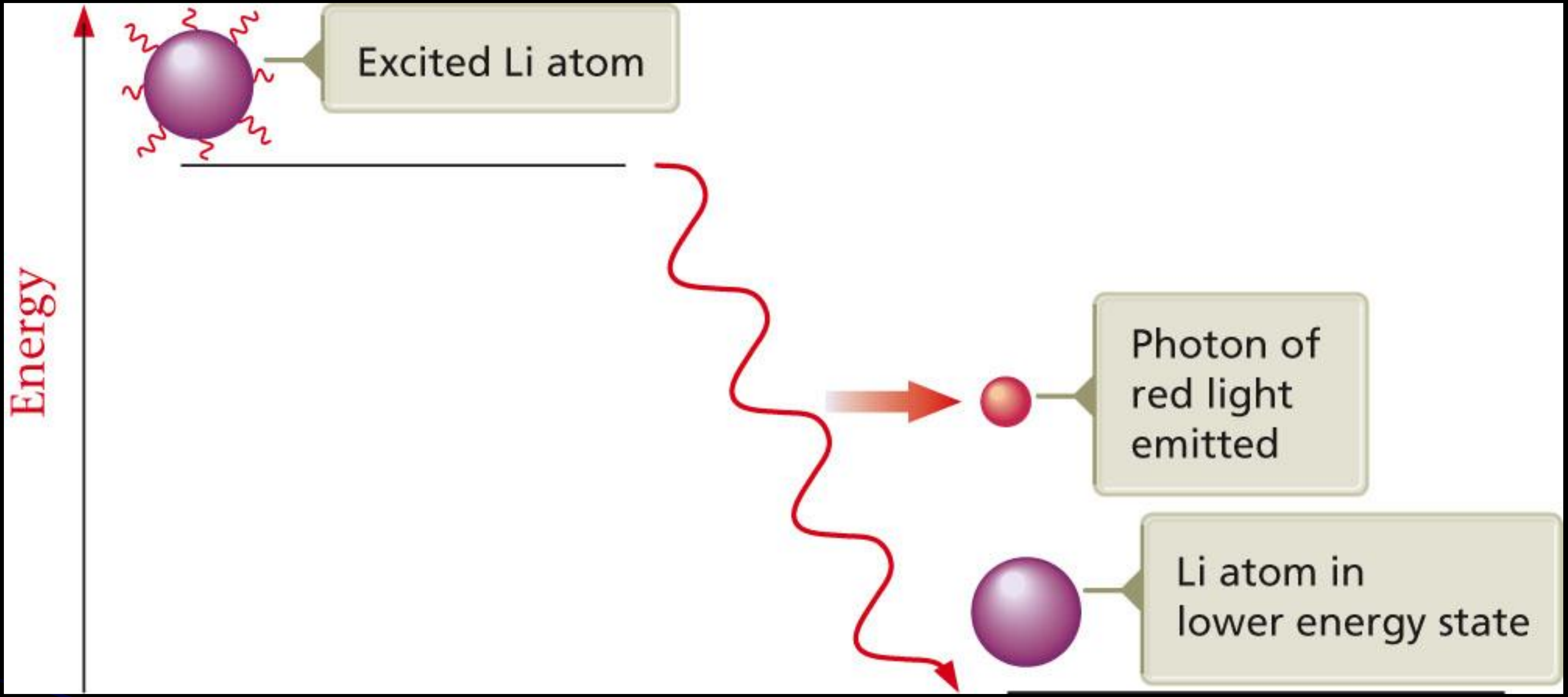
-The low energy state, after the excess energy is given off, is called the **ground state**.

-Different wavelengths of light carry different amounts of energy per photon.

Ex. Red light has lower-energy photons than blue light.

-After an atom has absorbed energy from some source, it uses this energy to enter and excited state.

-It releases this energy (goes to a lower energy state) by emitting photons of light.



An excited lithium atom emitting a photon of red light to drop to a lower energy state.

Hydrogen atomic emission spectrum

- Only certain wavelengths, not all wavelengths, are given off by hydrogen. Only certain energy changes are occurring.
- Thus hydrogen atoms have certain discrete energy levels. **Each energy level is quantized, meaning only certain values are allowed.**
- Scientists have found that all atoms are **quantized.**

Hydrogen atomic emission spectrum

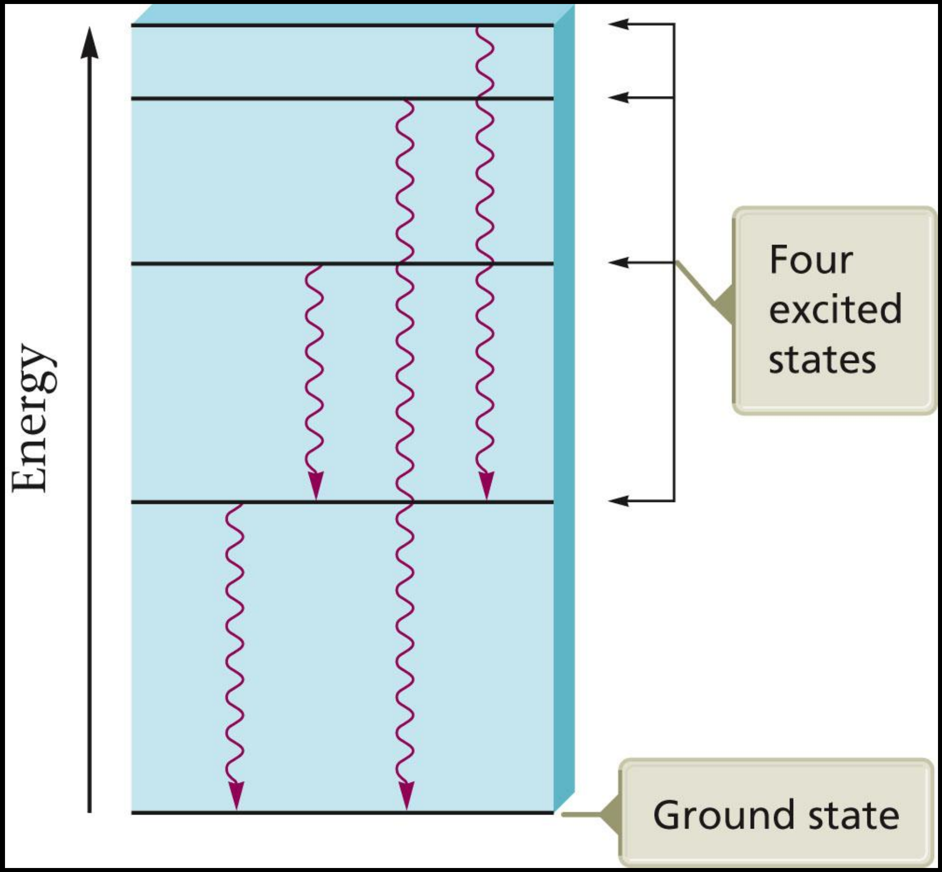
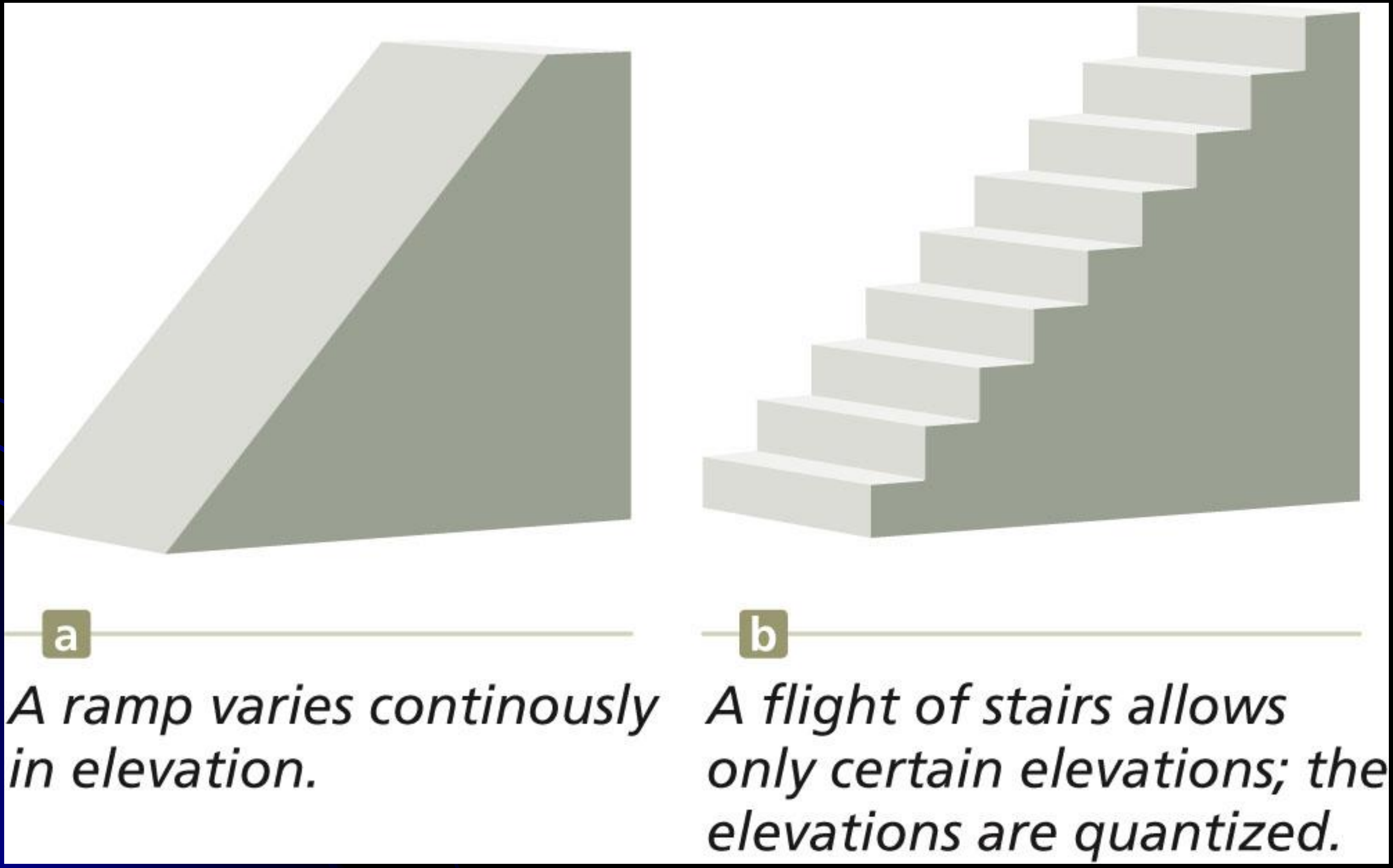
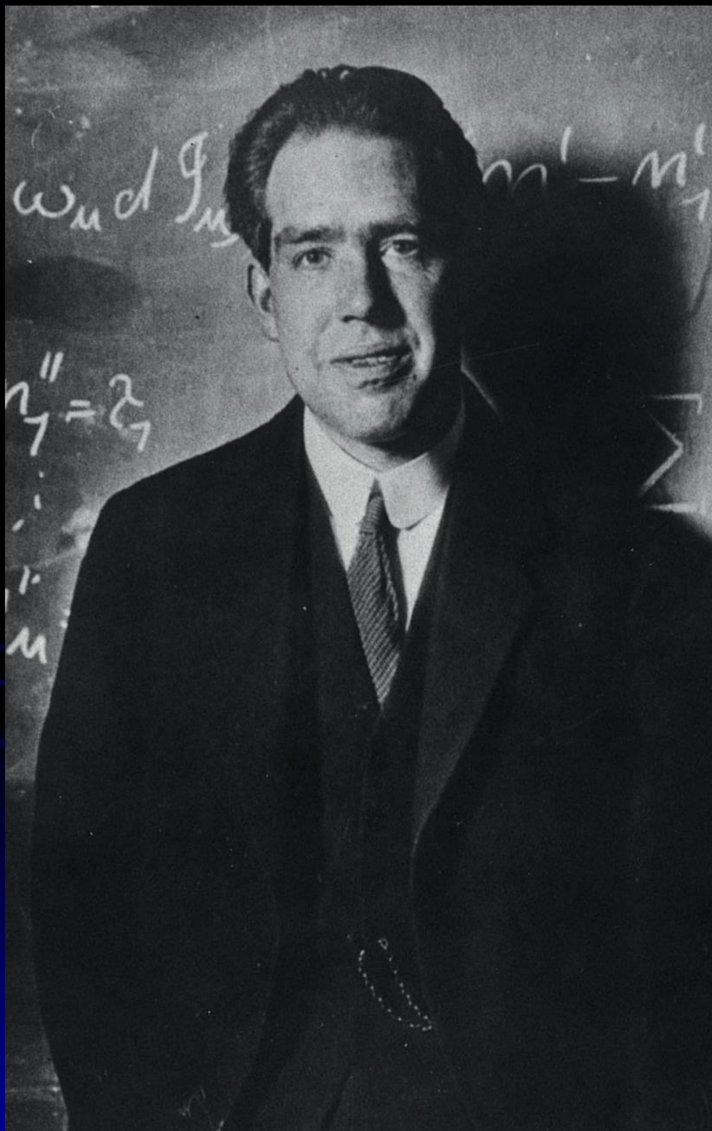


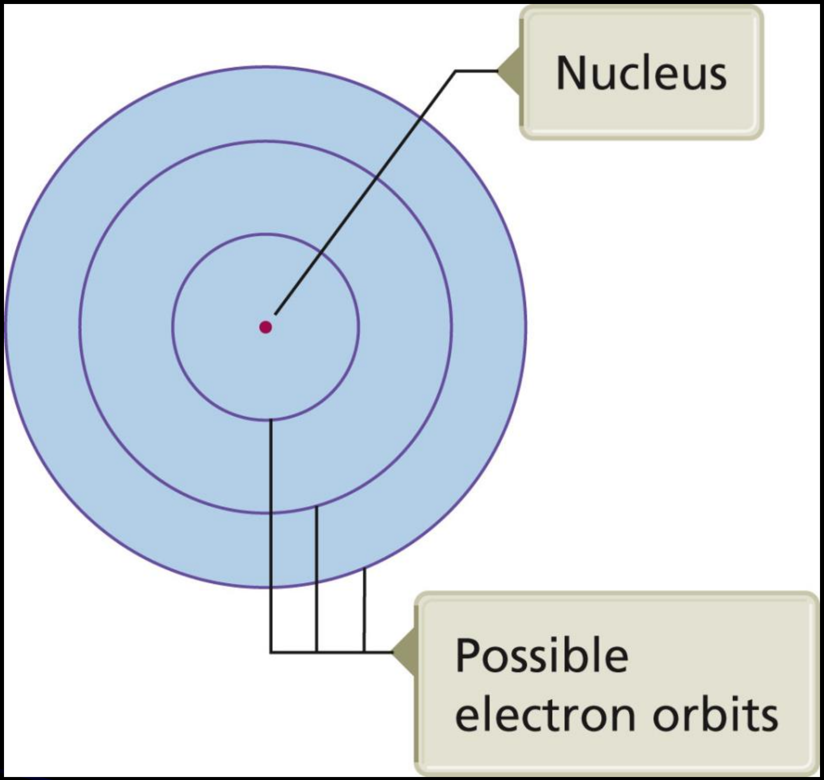
Figure 11.15: The difference between continuous and quantized energy levels can be illustrated by comparing a flight of stairs with a ramp.



11.5 The Bohr Model of the Atom



Bohr's model showed electrons traveling in different energy levels around the nucleus.



He predicted their movement to be circular orbits or paths.

-His model worked well for hydrogen but not other atoms.

-It has been found that the electron does not travel in circular orbits.

-Scientists do not exactly know how the electron moves.

11.6 The Wave Mechanical Model of the Atom

-A new model was needed.

-**Schrodinger** developed this model of the atom using mathematics.

-It describes the area where the electrons move in terms of **orbitals**, not orbits. Read pg. 332

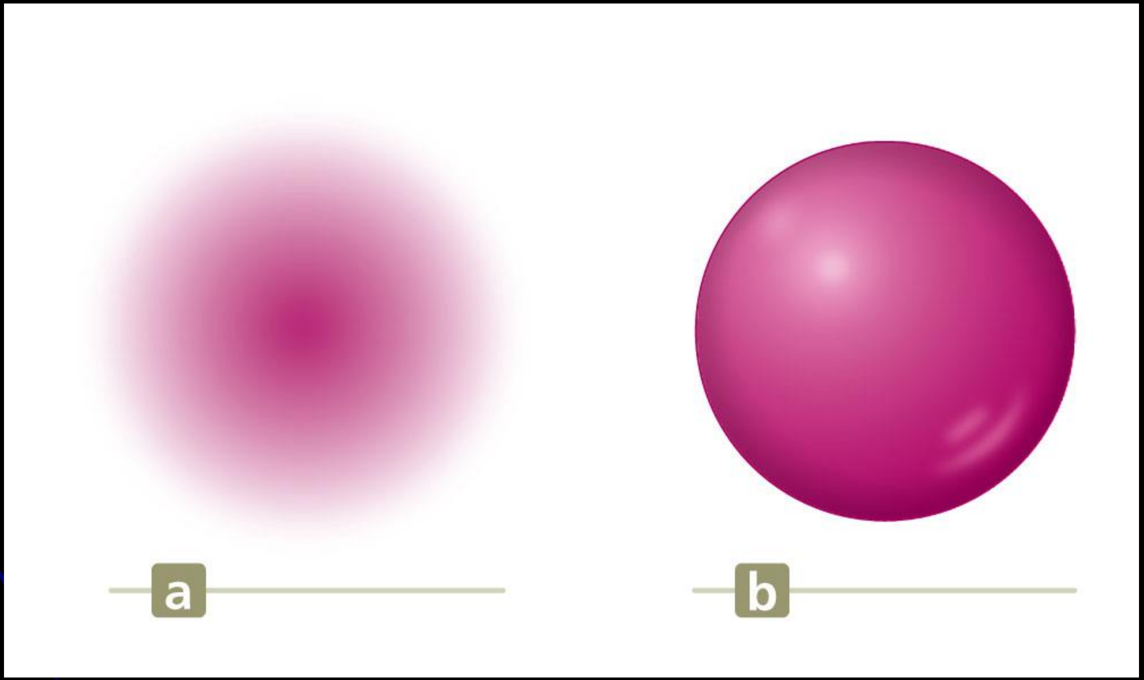
-**Orbitals** *are areas in space*, orbits are distinct paths.

-This model does not tell exactly where an electron is around the atom, but gives the *probability of finding an electron in a certain region*. (**Probability map**)

-This model tells nothing about when or how the electron moves.

11.7 The Hydrogen Orbitals

The orbital of an electron is the space around the atom in which the electron can be found 90% of the time.



-There are no sharp boundaries,

ex. atmosphere.

(*The electron does not travel on the surface of the orbital, it travels within the orbital*.)

-In the *ground state*, the electron in the hydrogen atom is found in an orbital called the **1s orbital**.

-If an electron absorbs energy it moves to a higher energy state and moves to another orbital. **Orbitals have different shapes**.

-Atoms have discrete **energy levels** called "**principal energy levels**."

These are labeled with integers, 1-∞.

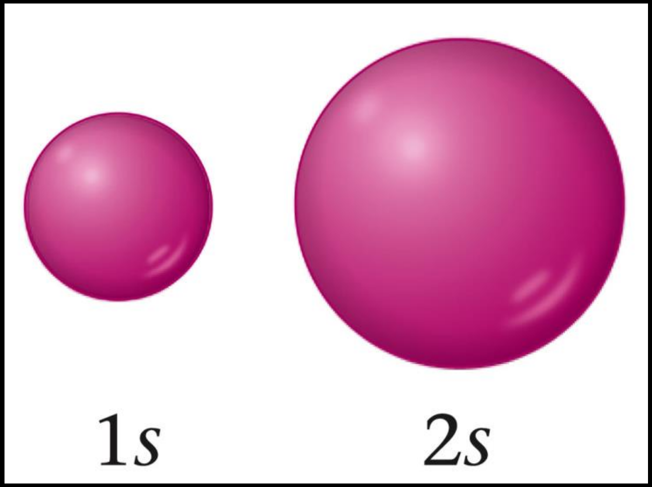
However, in our universe we only deal with 1-7.

-Each energy level consists of "**sublevels**."

Energy level	Sublevels
1	1
2	2
3	3
4	4
5	5 (4)
6	6 (3)
7	7 (2)

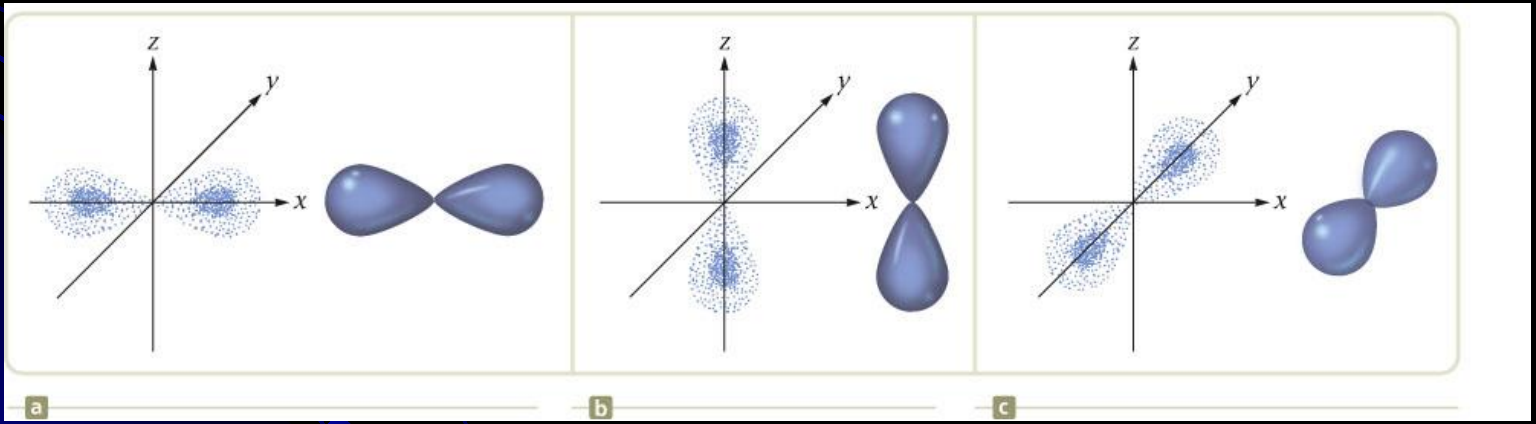
We will study these four types of sublevels:

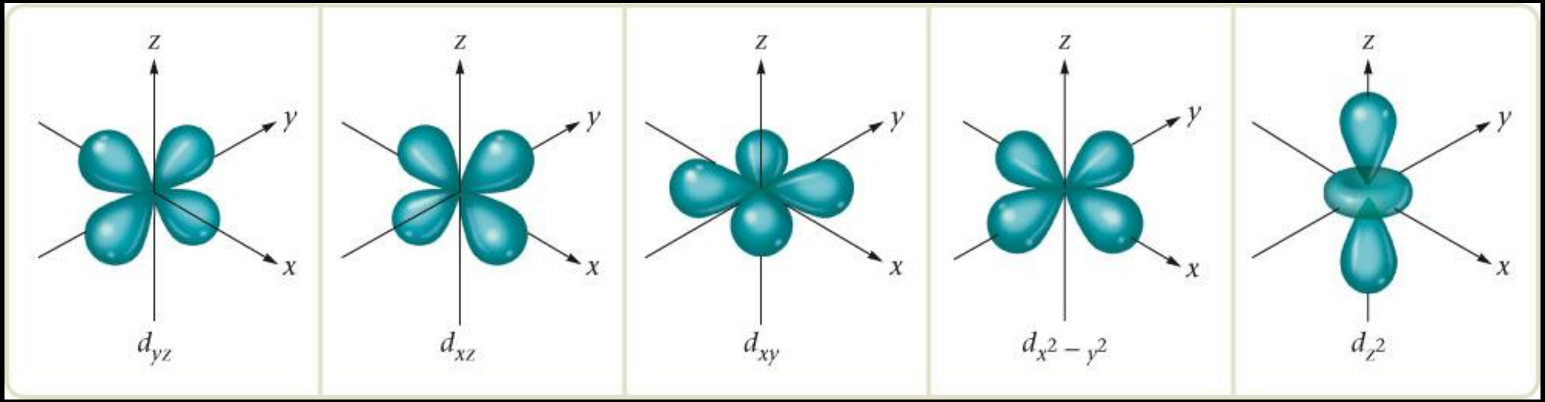
s, p, d, f



S is spherical in shape

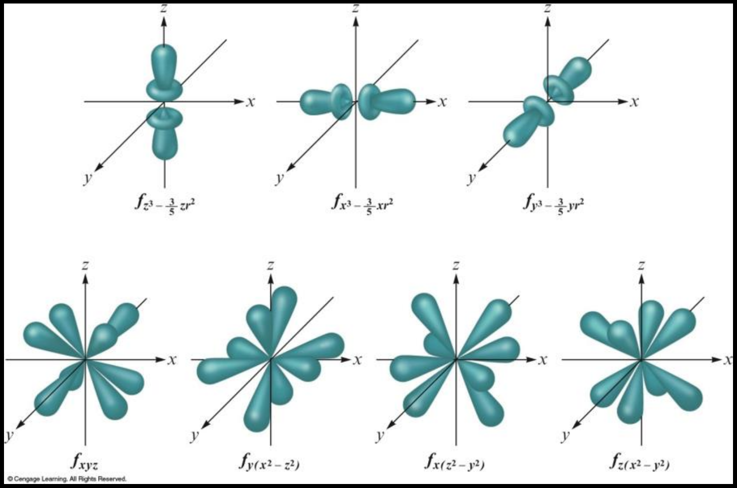
p is dumbbell shaped



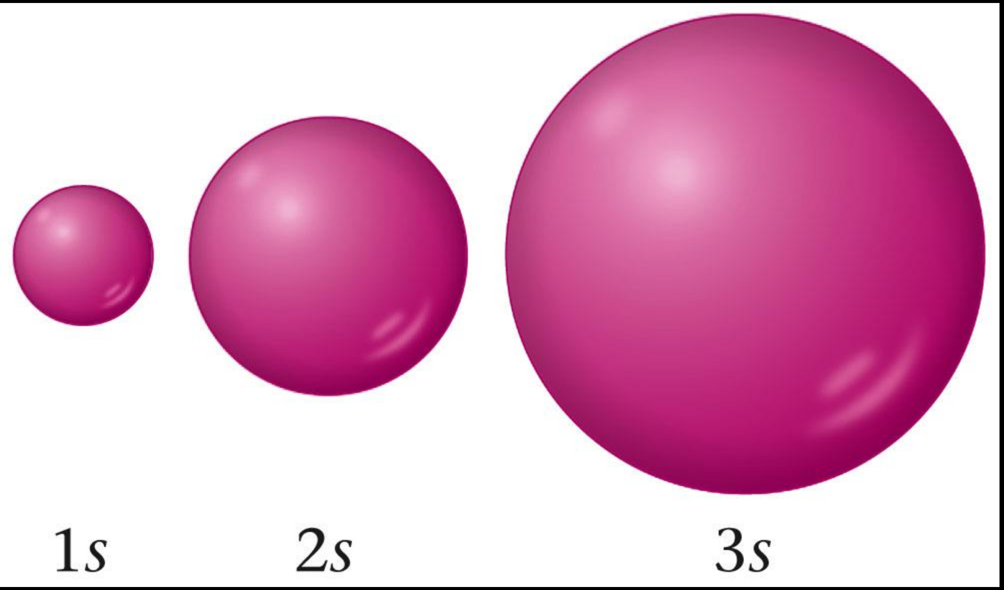


d **is flower shaped**

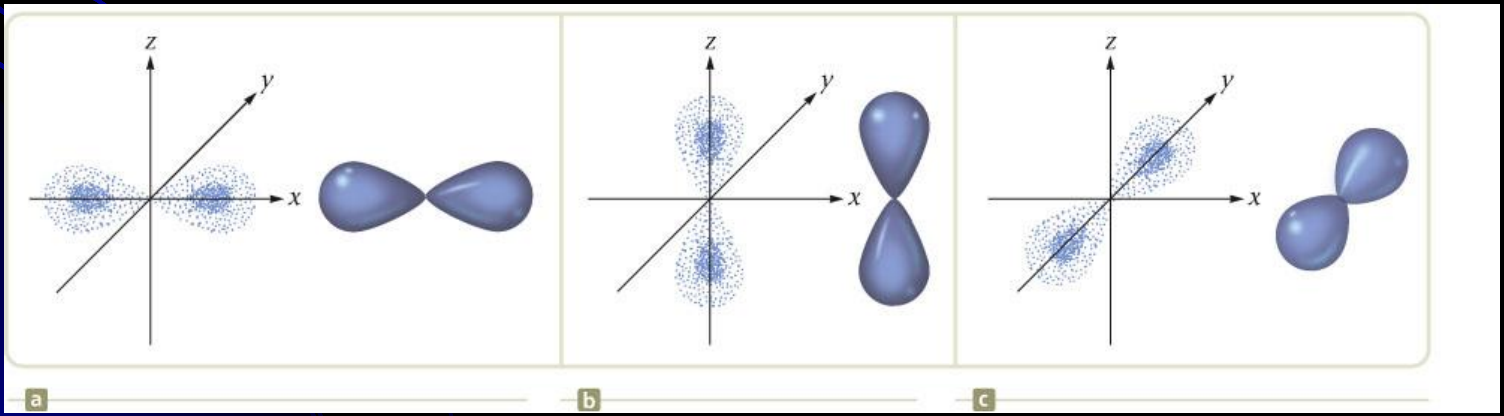
f **is very complex**

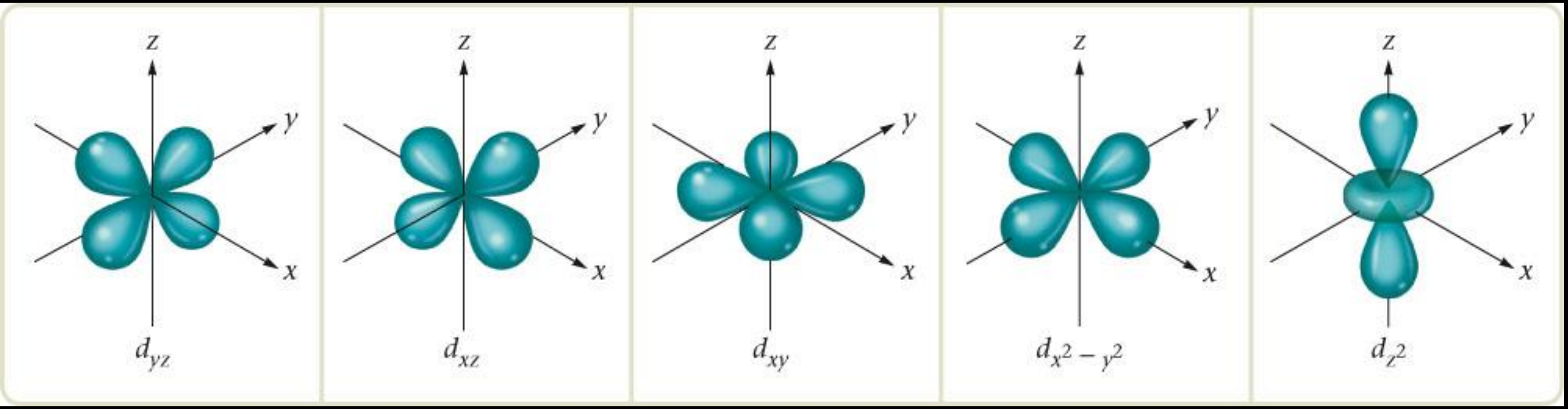


- Sublevels contain spaces called "**orbitals**"



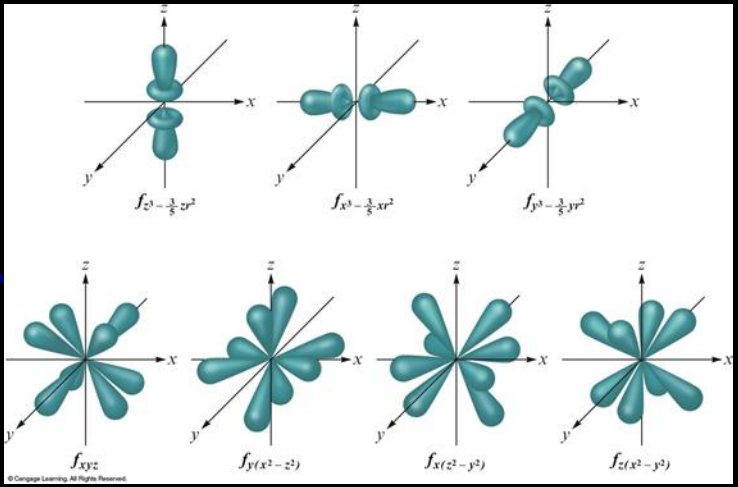
- Each sublevel has a different number of orbitals.
  - **s = 1 orbital**
  - **p = 3 orbitals**





d = 5 orbitals

f = 7 orbitals



Energy Level	Sublevel	Orbitals
1	s	1
2	s, p	4
3	s, p, d	9
4	s, p, d, f	16
5	s, p, d, f	16
6	s, p, d	9
7	s, p	4

11.8 The Wave Mechanical Model: Further Development

-Each orbital can contain up to **two electrons**.

-If the orbital has two electrons, they spin in **opposite directions**.

(Spin is represented by arrows: ↑ and ↓)

1) **Pauli Exclusion Principle**: *an orbital can hold at maximum two electrons.*

Indicate whether each of the following statements about atomic structure is true or false.

a. An *s* orbital is always spherical in shape.

b. The 2*s* orbital is the same size as the 3*s* orbital.

c. The number of lobes on a *p* orbital increases as n increases. That is, a 3*p* orbital has more lobes than a 2*p* orbital.

d. Level 1 has one *s* orbital, level 2 has two *s* orbitals, level 3 has three *s* orbitals and so on.

e. The electron path is indicated by the surface of the orbital.

2) **Aufbau Principle**: electrons enter the orbitals of lowest energy first.

3) **Hund's Rule:** when electrons occupy orbitals of equal energy, one electron enters each orbital until all orbitals contain one electron with parallel spins.

Each orbital in a sublevel has equal energy.(**degenerate**)

11.9 Electron Arrangements for Elements 1-18

For the first 18 elements the electrons fill the sublevels in this order: 1*s* 2*s* 2*p* 3*s* 3*p*

Hydrogen: 1 electron → 1*s*

Electron configuration

$1s^1$

Orbital(box) Diagram

$\boxed{\uparrow}$

1*s*

Helium: 2 electrons $\rightarrow$ 1*s*

Electron configuration

$1s^2$

Orbital(box) Diagram

$\boxed{\uparrow\downarrow}$

$1s$

Lithium: 3 electrons $\rightarrow$

$1s^2 2s^1$

$\underset{1s}{\boxed{\uparrow\downarrow}}\quad\underset{2s}{\boxed{\uparrow}}$

Carbon: 6 electrons $\rightarrow$

$1s^2 2s^2 2p^2$

$\underset{1s}{\boxed{\uparrow\downarrow}}\quad\underset{2s}{\boxed{\uparrow\downarrow}}\quad\underset{2p}{\boxed{\uparrow}\boxed{\uparrow}\boxed{}}$

H $1s^1$							He $1s^2$
Li $2s^1$	Be $2s^2$	B $2p^1$	C $2p^2$	N $2p^3$	O $2p^4$	F $2p^5$	Ne $2p^6$
Na $3s^1$	Mg $3s^2$	Al $3p^1$	Si $3p^2$	P $3p^3$	S $3p^4$	Cl $3p^5$	Ar $3p^6$

The electron configurations in the sublevel last occupied for the first eighteen elements.

Compare: **electron configurations of elements in the same group**

Ex. H, Li, and Na

H = $1s^1$

Li = $1s^2 2s^1$

Na = $1s^2 2s^2 2p^6 3s^1$

Notice the highest energy level contains the same number of electrons. These are the ***valence*** *electrons*. **The valence electrons are the ones involved in chemical reactions.**

Since elements in the same group have the same number of valence electrons, the chemical properties are very similar.

11.10 Electron Configurations and the Periodic Table

For elements greater than 18 the pattern is different. You might expect: 1s2s2p3s3p3d4s4p4d4f5s5p5d etc.

But really is:

1s2s2p3s3p4s3d4p5s4d5p6s5$d^1$4f5$d^9$6p7s6d^1 5f6d^9

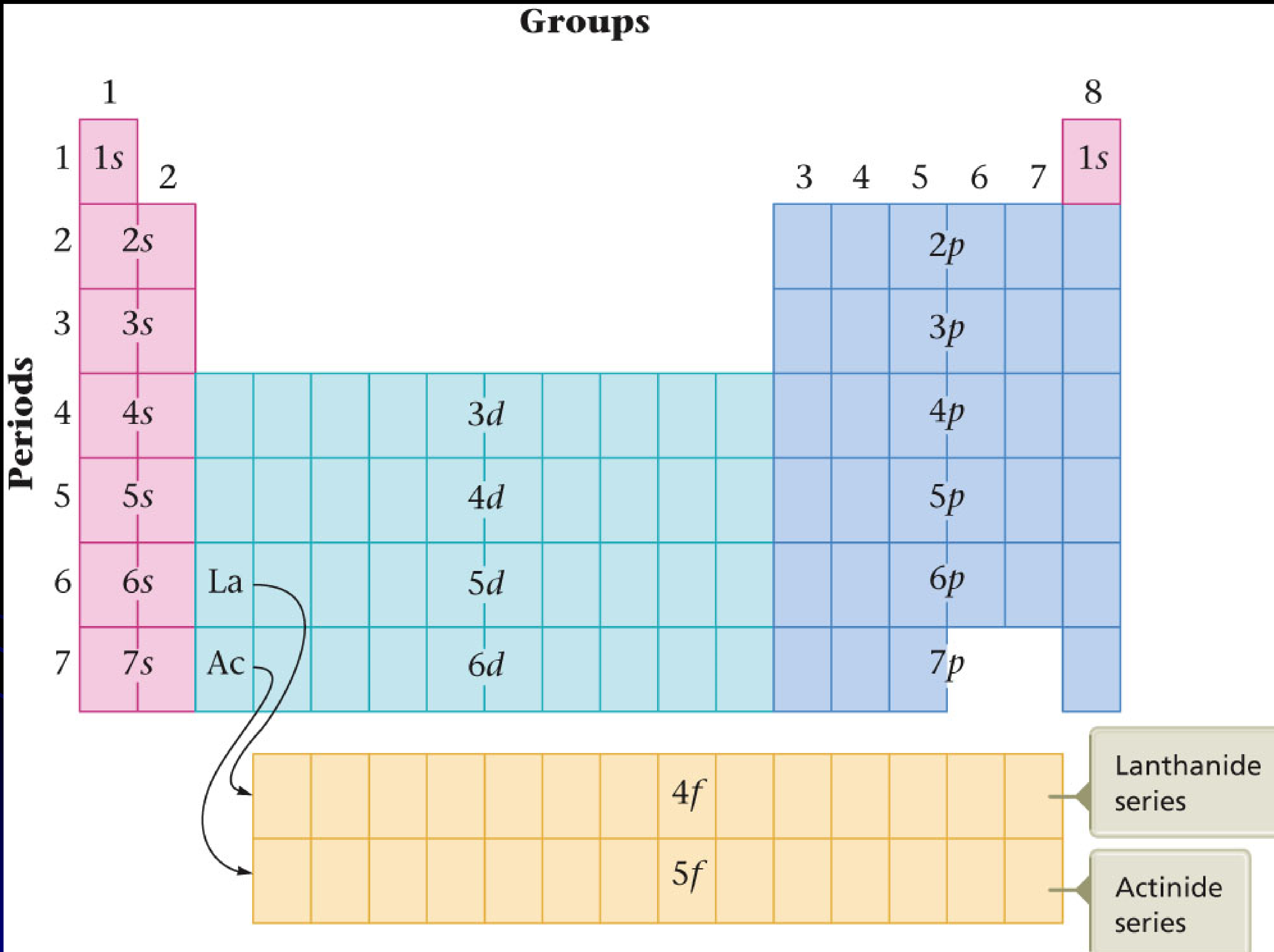
Don't memorize this. Use the periodic table.

Column 1 and 2 = ***s***

Column 3-8 = ***p***

Transition elements = ***d***

Inner Transition Elements = **f**


Groups
Periods
1
2
3
4
5
6
7
8
1s
2s
3s
4s
5s
6s
7s
2p
3p
4p
5p
6p
7p
3d
4d
5d
6d
La
Ac
4f
5f
Lanthanide series
Actinide series

Each row represents the energy level.

Start from the top and read left to right.

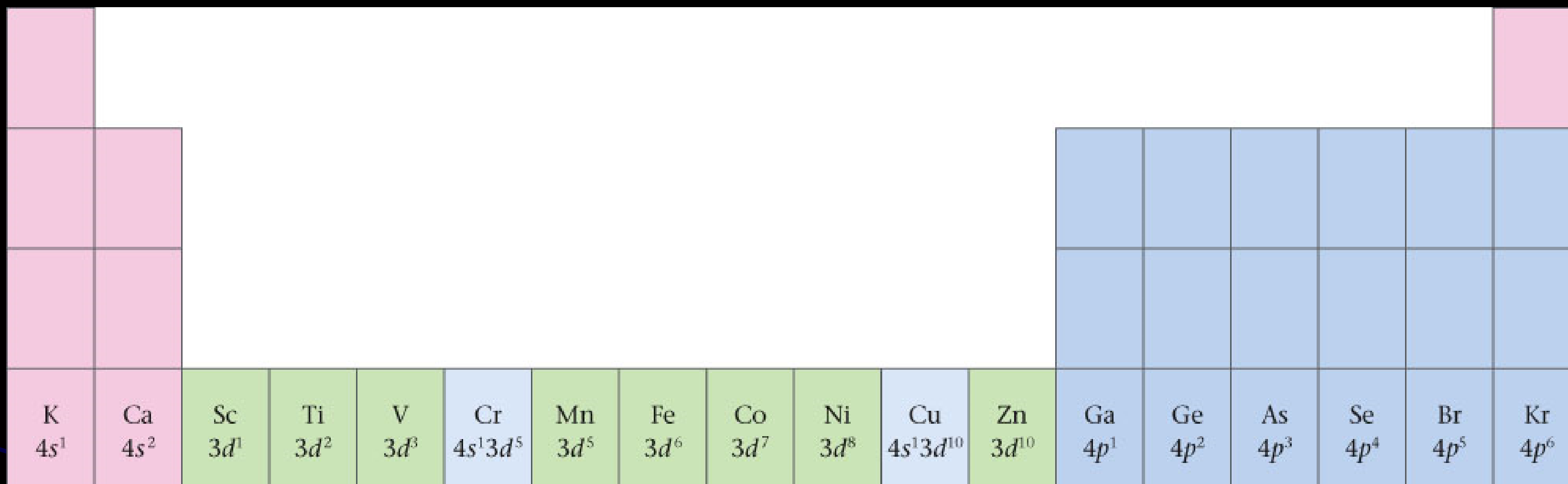
Group d subtract 1 from the row number

Group f subtract 2 from the row number.

Fe $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$

Zr $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^2$

Some elements have exceptional electron configurations.



Don’t memorize these but know they exist.

*Know these terms: Lanthanide, actinide, representative elements.

Shortcut method:

Instead of writing the inner electrons;

Na = **$1s^2 2s^2 2p^6$** **$3s^1$**

$1s^2 2s^2 2p^6$ is the electron configuration of **Ne** therefore we can write [Ne]$3s^1$.

The shortcut for the **noble gases** is just the symbol in brackets.

11.11 Atomic Properties and the Periodic Table

Science is based on observation. *Why study atomic theory?* **To help us better understand our world and how it works.**

Read pg. 347-351

-most reactive metals are found lower-left on the periodic table.

-electrons are removed easily because of the distance from the nucleus.

-most reactive nonmetals are found upper-right on the periodic table.

-these elements pull electrons from metals very effectively because of the large positive charge in the nucleus.

-**Ionization energy**: *is the energy required to remove an electron from an individual atom in the gas phase.* (Metals have low ionization energies)

Trend: ***across the period*** ionization energy **increases**

down a group ionization energy decreases

-**Atomic size**:

Trend:

across the period

atomic size decreases

down a group

atomic size increases