

**4 • Electronic Structure & the Per. Table  
Wave Ideas You Should Know  
(1 of 16)**

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<b>EMR</b>	electromagnetic radiation ... oscillating electric & magnetic fields at right angles
<b>wavelength ( <math>\lambda</math> )</b>	the distance from crest to crest or trough to trough.
<b>amplitude</b>	the distance from the equilibrium point to the crest or trough.
<b>frequency ( <math>\nu</math> )</b>	the number of waves that pass a point per second. (Hz, $s^{-1}$ , $1/s$ )
<b>continuous spectrum</b>	a “normal” rainbow that contains all of the colors (ROYGBV).
<b>line spectrum</b>	a spectrum that only contains certain bright lines that result from electron transitions within an atom.

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**4 • Electronic Structure & the Per. Table  
Wave Calculations  
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**Exchanging wavelength, frequency, & energy of light:**

$$\lambda \nu = c \quad c = 3.00 \times 10^8 \text{ m/s [speed of light]}$$

$$E = h \nu \quad h = 6.626 \times 10^{-34} \text{ J}\cdot\text{s [Planck's constant]}$$

$$E = \frac{hc}{\lambda}$$

**Energy of Level “n” in the Hydrogen atom:**

$$E_n = -\frac{A}{n^2} \quad A = 2.18 \times 10^{-18} \text{ J [Arrhenius constant?]}$$

NOTE: You can learn the Balmer equation or the Rydberg equation listed on pp. 108 & 109. I suggest you use the above equations to calculate the wavelengths or frequencies of light emitted when electrons change energy levels.

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**4 • Electronic Structure & the Per. Table  
The Balmer Series  
(3 of 16)**

When the electron “drops” to energy level  $n=2$ , visible light is emitted. The bright line spectrum observed is called the **Balmer Series** of lines. [Memorize this info...]

3	2	red
4	2	blue-green
5	2	blue-violet
6	2	violet

All the lines result from an electron transition to level  $n=1$  are too high energy to be visible... UV [Lyman Series]

All the lines that result from an electron transition to levels  $n=3$  [Paschen],  $n=4$  [Brackett] and  $n=5$  [Pfund] are too LOW energy to be visible.

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**4 • Electronic Structure & the Per. Table  
The De Broglie Wavelength of Electrons  
(4 of 16)**

Two equations can be combined into one to allow you to calculate the **wavelength** of a **particle**:

$$E = mc^2 \quad E = \frac{hc}{\lambda}$$

$$mc^2 = \frac{hc}{\lambda}$$

$$mc = \frac{h}{\lambda} \quad \text{or more generally, } mv = \frac{h}{\lambda}$$

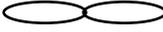
$$= \frac{h}{mv}$$

$h$  = Planck's constant,  $m$  = mass of particle (in kg),  
 $v$  = velocity of particle (in m/s).

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4 • Electronic Structure & the Per. Table  
Standing Waves  
(5 of 16)

**Standing waves** are something that waves do...  
resulting from having repeating waves that always cancel at  
the nodes and always add up at the antinodes.

We have seen standing waves  
of strings: 

of sound: ||| | | | | |||| | | | | | |||

of drumheads: 

We hypothesized the standing waves  
of electrons: (4-dimensional vibrating wavicles)  
the orbitals (s, p, d, f... probability waves)

4 • Electronic Structure & the Per. Table  
Quantum Numbers (n, l, m, s)  
(6 of 16)

These are **variables** in some unseen equation:

**The Rules:**

$n = 1, 2, 3, \dots$  determines the energy of the  $e^-$

$l = 0 \quad (n-1)$  the type of orbital (subshell)  
0 s, 1 p, 2 d, 3 f

$m = -l \quad +l$  which orientation of the orbital  
(x, y, z... for p orbitals)

$s = +\frac{1}{2}$  or  $-\frac{1}{2}$  the "spin" of the electron

NOTE: m is also called  $m_l$  and s is  $m_s$

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Three Rules for Filling Orbitals  
(7 of 16)

The **Pauli Exclusion Principle** states that no two electrons  
in an atom may have the same four quantum numbers. This  
translates to the idea that an orbital may contain no more  
than two electrons.

The **Aufbau Principle** states that electrons occupy the  
lowest energy available orbital. You must memorize the  
orbital chart on page 120. [Aufbau = "building up"]

**Hund's Rule** states that when you have several orbitals of  
the same energy (e.g. three p orbitals or five d orbitals) place  
one electron in each orbital before doubling them up.

NOTE: Remember the Aufbau hotel analogy...

4 • Electronic Structure & the Per. Table  
Electron Configurations  
(8 of 16)

This is a **shorthand** notation for the arrangement of electron  
in the orbitals.  $1s^2$  means 2 electrons in the 1s orbital

Consider Arsenic, As

This is the order in which the electrons fill...

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

**however**, we write the orbitals according to the distance  
from the nucleus ( $n=3$ 's then  $n=4$ 's, etc.)

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

long form: (shown above) show each subshell

short form: (show the last filled NRG level...noble gas core)

As [Ar]  $3d^{10} 4s^2 4p^3$

#### 4 • Electronic Structure & the Per. Table The s-block, p-block... of the Periodic Table Exceptions to the Filling Rules

(9 of 16)

#### 4 • Electronic Structure & the Per. Table Shapes of the orbitals (10 of 16)

#### 4 • Electronic Structure & the Per. Table Predicting the Atomic Size (radius) Trends in the Periodic Table (11 of 16)

#### 4 • Electronic Structure & the Per. Table Explaining The Sizes of Ions The Lanthanide Contraction (12 of 16)

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The shape of the Periodic Table comes from the way the electrons fill the orbitals.

s-block	Families I and II
p-block	Families III, IV, V, VI, VII, and VIII
d-block	The Transition Elements
f-block	The Inner Transition Elements

Six elements do not follow the rules. They are in the d-block of the periodic table... the transition elements.

Cu, Ag, and Au... instead of a full s-orbital and almost filled d-orbital, they have a filled d and half-filled s-orbital  
Cr, Mo, W... instead of a full s-orbital and an almost half-filled d-orbital, they have  $d^5$  and  $s^1$

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Know the general shapes of the s (spherical) orbitals,



p (perpendicular) orbitals, and



d (diagonal) orbitals.



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What: As you move ACROSS a period, the size of the atom DECREASES

Why: As you move ACROSS a period, the number of protons in the nucleus increases... so the proton-electron attraction increases... the size decreases.

NOTE: The increased  $e^-$ - $e^-$  repulsion that one might expect is not important, because the outer electrons do not SHIELD the electrons from the nucleus..

What: As you move DOWN a family, the size of the atom INCREASES

Why: As the value of "n" increases, the average distance of the electron from the nucleus increases. The size of the atom IS the electron cloud... so the size increases

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The change in size of ions (compared to the neutral element) depends on the electron-electron repulsion.

If an atom gains electrons, the increased repulsion increases the size of the electron cloud. So... negative ions are larger than the neutral atom.

Positive ions form by losing electrons (less repulsion) and get smaller.

The Lanthanide contraction makes the third row of the transition elements about the same size as the second row... This causes these elements to be much more dense.

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**4 • Electronic Structure & the Per. Table  
Ionization Energy  
Trends Across a Period  
(13 of 16)**

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Following the decrease in size, the ionization energy **GENERALLY** increases as you move across a period. If the atom is smaller, the electron being removed is closer to the nucleus and therefore feels a stronger attraction.

The Jags: (why are some electrons **EASIER** to remove)  
In Family III, the electron being removed comes from the p-orbital rather than the s-orbital. The p-orbital electron is at a higher energy and requires **LESS** energy to ionize.

In Family VI, the electron being removed comes from an orbital with **TWO** electrons. The  $e^- - e^-$  repulsion felt by the electron allows it to be ionized with less added energy.

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**4 • Electronic Structure & the Per. Table  
Ionization Energy & Reactivity  
Trends Down a Family  
(14 of 16)**

What: Down a family, Ionization energy **DECREASES**  
Why: The trend here goes right with the size of the atom. Since “n” increases, the distance of an electron from the nucleus increases and is easier to remove.

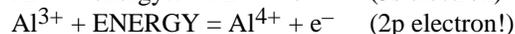
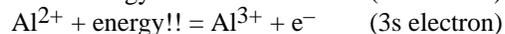
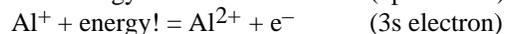
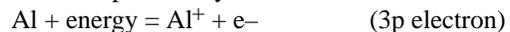
What: Down family I and II, the reactivity **INCREASES**.  
Why: These families lose electrons to become + ions. The easier it is (lower ionization energy) the more reactive they are.

What: Down family VII, the reactivity **DECREASES**.  
Why: These elements **GAIN** electrons to become – ions. Smaller atoms mean the attracted electrons will be closer to the nucleus... more effective attractions.

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**4 • Electronic Structure & the Per. Table  
Ionization Energy  
Trends in Successive Ionizations  
(15 of 16)**

What: You can tell the family of an element by observing when the ionization energies get very large.  
For example: family III elements

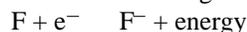


Why: The “easy” electrons to remove are in an orbital with a higher value of n. When n decreases, the average distance of the electron from the nucleus decreases and the attractions between the proton and electrons increase.

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**4 • Electronic Structure & the Per. Table  
Electron Affinity  
(16 of 16)**

Electron affinity is the energy involved when an atom gains an electron to become a negative ion.



Elements in the upper right corner of the periodic table have the greatest electron affinity (greatest attraction for electrons).

The electron affinity may be + or –. Negative values mean energy is released and also counts as a greater electron affinity.

Electron affinity data is not complete, but it gives **SOME** evidence for trends in the periodic table.

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