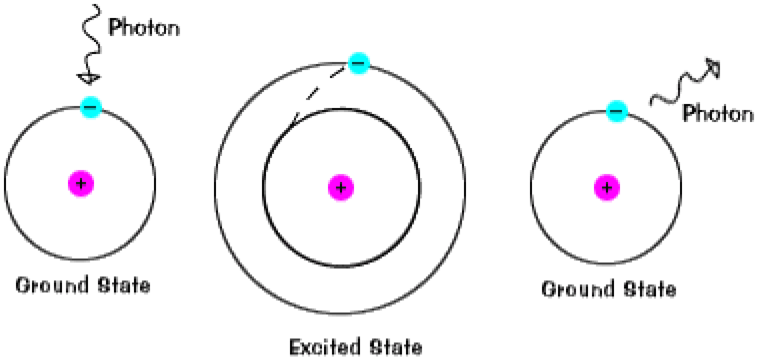
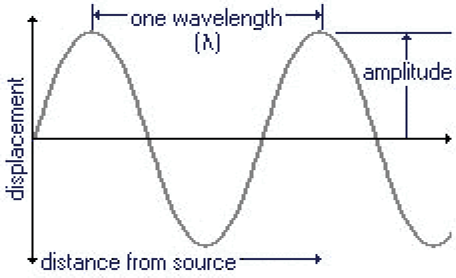
Notes: Chapter 11

- The source of all light is the atom
- Light is generated when the electrons in an atom are excited to a higher energy level and then relax. When the atoms relax, they emit photons as the electrons drop to lower energy levels:



  - Therefore, the more we understand about light, the more we can understand about the arrangement of electrons in the atom.
  - This last point is important because the electron is the most important subatomic particle in chemistry. The arrangement of an atom’s electrons, its **electron configuration, is what determines an element’s chemical properties.**
  - We will first discuss light as a wave phenomenon. Later, we will describe light in terms of being a particle.
  - Light is a transverse wave. A transverse wave is one in which the medium travels at right angles to the direction of the wave energy that is traveling through it.
  - A wave is simply some disturbance in a medium
  - A good example of a transverse wave is an ocean wave. The medium in this case would be the ocean.

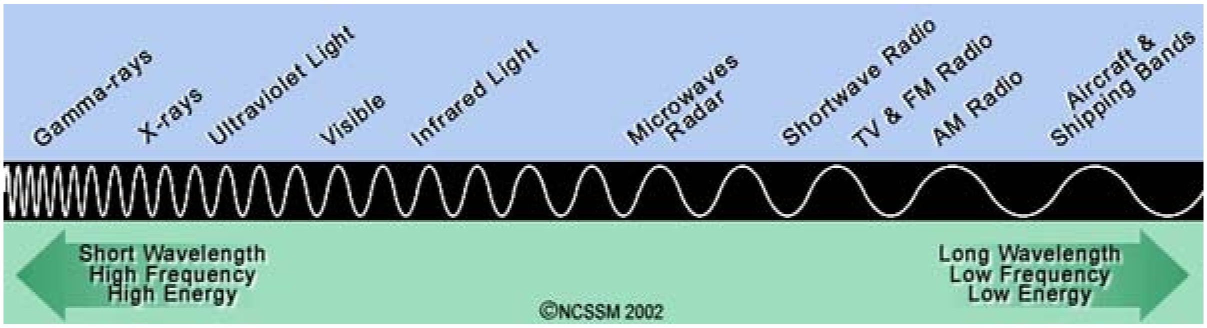


- You should understand these terms used to describe waves:
- Wavelength
  - Symbol: l (pronounced and written in English “lambda”)
  - Measured in meters
  - Equals the distance between two crests (top of wave) or troughs (the bottom of the wave)
- Frequency
  - Symbol: n (pronounced and written in English “nu”)
  - Measured in Hertz (Hz), or s^{-1}, or 1/s. The latter two are called “reciprocal seconds”
  - Equals the number of wave cycles that pass a given point per given amount of time.
- Speed
  - Symbol is c in the case of light, which is the only case we are concerned with

  - Is typically measured in meters per second (m/s)
  - The speed of light is ALWAYS equal to 300,000,000 m/s in a vacuum.
  - c =3.00 x 10^8 m/s
  - Because this is a speed that never changes, the speed of light is called a constant.
  - Speed = frequency X wavelength (for any wave)
  - For light,

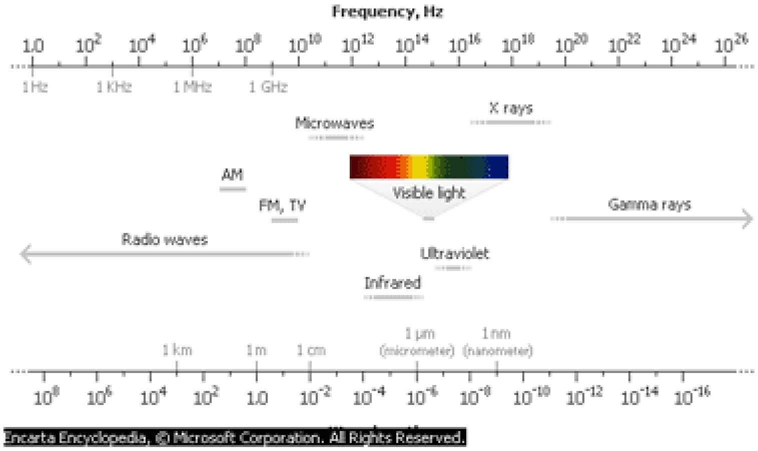
$$c = \mathrm{ln}$$

- From this equation, and from the demonstrations in class, we can see that
  - Frequency and wavelength are **inversely related**
  - Frequency and energy are **directly related**
  - We will see that this is true of the different types of electromagnetic radiation (light) as well.



- Three examples of speed/wavelength/frequency problems
  - What is the frequency of light that has a wavelength of 250 m?

- Three examples of speed/wavelength/frequency problems
  - What is the frequency of light that has a wavelength of 250 m?
  - All light is called electromagnetic energy. These terms (light and E.R.) are synonymous for our purposes.



Electromagnetic Spectrum

The electromagnetic spectrum includes radio waves, microwaves, infrared light, visible light, ultraviolet light, x rays, and gamma rays. Visible light, which makes up only a tiny fraction of the electromagnetic spectrum, is the only electromagnetic radiation that humans can perceive with their eyes.

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  - Only a small portion of the electromagnetic spectrum is visible light. That is, of all of the light that exists, only part of it can be seen by humans.

- Visible light spectrum (p. 299)

The visible light spectrum. From left to right, in order of decreasing energy, decreasing frequency, and increasing wavelength.

- The electromagnetic spectrum, in order of increasing frequency, increasing energy, and decreasing wavelength:
    - Radio waves
    - Microwaves
    - Infrared (IR) radiation
    - Visible light
    - Ultraviolet light
    - X-rays
    - Gamma rays
- The visible light spectrum, in order of increasing energy, increasing frequency, and decreasing wavelength:
    - Red
    - Orange
    - Yellow
    - Green
    - Blue
    - Indigo
    - Violet
- Practice problems:
    - What is the wavelength of light that has a frequency of 1.2 x 10^{15} Hz?
    - What is the frequency of light that has a wavelength of 650 nm (10^9 nanometers = 1m)? According to the diagram on p. 299, is this light visible? How did you know?

- What is the wavelength of light that has a frequency of 1.2 x 10^{15} Hz?
- What is the frequency of light that has a wavelength of 650 nm? (10^9 nanometers = 1m) According to the diagram on p. 299, is this light visible? How did you know?

- Light can also be thought of as a particle.
- The quantum theory was first suggested by Max Planck in 1900. He suggested that energy can only be absorbed or emitted in small packets called quanta (singular: quantum).
- Money can only be spent in tiny amounts called cents. Every bit of money that you have ever spent or earned in the U.S. has been some multiple of 1 cent.
- **Energy is either absorbed or emitted (gained or lost) in multiples of *h*, Planck's constant. $h = 6.626 \times 10^{-34}$ Js**
- A quantum of light energy is called a **photon**. **A photon is light a tiny packet of light energy.**
- To calculate the energy of a photon of light, we use the formula

$$E = hn$$

  - E = energy of a photon
  - h = Planck's constant = 6.626×10^{-34} Js
  - n = frequency of light

- Practice problems:
  - Calculate the energy of a photon that has a frequency of 1.20×10^{15} Hz.
  - Calculate the frequency of a photon with an energy of 3.90×10^{-19} J.
- Do problems 1-6, ch. 11, in class.

- Do worksheet in class

General Chemistry
Mr. MacGillivray
Worksheet:
Light and Quantum Theory

1. Arrange the seven types of electromagnetic radiation that we discussed in class in order of **DECREASING** energy:

a. ________ (highest E)
b. ________
c. ________
d. ________
e. ________
f. ________
g. ________ (lowest E)

2. In the list above, use words and arrows to indicate how the wavelength and frequency are changing.

3. Repeat #1 and #2 with the colors of the visible spectrum.

  a. _______ (highest E)
  b. _______
  c. _______
  d. _______
  e. _______
  f. _______
  g. _______ (lowest E)

4. “If the wavelength of light is very short, then the energy is very __________ and the frequency is very _____________.”

5. “If the wavelength of light is very long, then the energy is very __________ and the frequency is very _____________.”

6. Wavelength and frequency are __________ly related. Energy and frequency are ___________ly related.

7. Energy is measured in these units: ________.
8. Wavelength is measured in these units: ________.
9. Frequency is measured in these units: ________, also written as ______ or ______.
10. Convert the following wavelengths to nm:
    a. l = 513 m

    b. l = 8.03×10^{-6} m

11. Convert the following wavelengths to m:
    a. l = 755 nm

    b. l = 0.272 nm

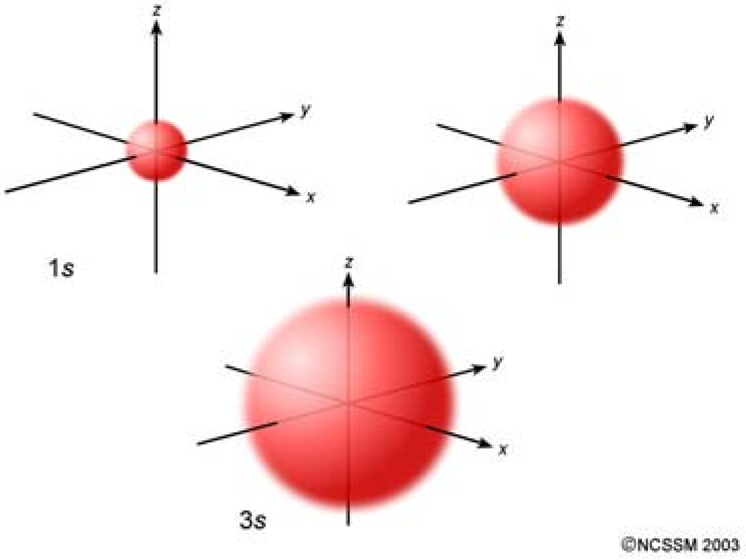
12. Find the energy of a photon of light with a frequency of 5.22×10^{21} 1/s.

13. Find the energy of a photon of light with a wavelength of 425 nm.

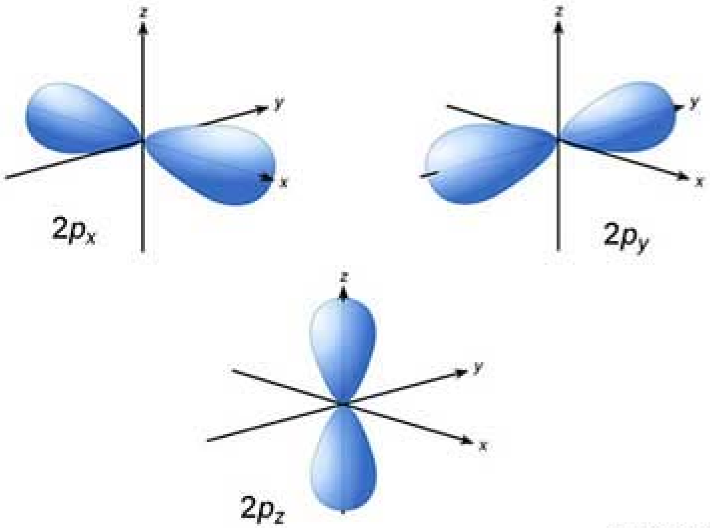
14. Find the wavelength of light with a frequency of 5.28×10^{15} s^{-1}.

15. Using p. 299, answer these questions:
    a. Is the light in question #12 visible?
    b. How did you know?
    c. Is it too high in energy or too low in energy to be seen?
    d. What type of light is it (what region of the electromagnetic spectrum)?

- Begin electron configurations
- Chemistry: it’s the study of matter and the changes it undergoes
- These changes are due to atoms combining, separating, or rearranging.
- These changes occur when electrons are given, taken, or shared between atoms
- The more we know about the electronic structure of the atoms,
- There are 3 levels of organization to the electron cloud.
    - Energy level (“n”)
    - Energy sublevel (“l”)
    - Orbital (“m_l”). An orbital is a region in space where electrons are likely to be found. It can contain at most two electrons.
- Think of the atom as a hotel with different floors (energy levels), rooms on each floor (energy sublevels), and closet/bathroom/bedroom/etc. (orbital) within each hotel room.
- Each energy level has a different number of sublevels
    - The first energy level has 1 sublevel, “1s”
    - The second has energy level has 2 sublevels, 2s and 2p
    - The third energy level has three energy levels, 3s, 3p, and 3d
    - The fourth energy level has four energy levels, 4s, 4p, 4d, and 4f
    - The fifth has five, and so on.
    - How many sublevels does the 7th energy level have?
- Sublevels are in turn divided up into **orbitals, which are regions in space where electrons are likely to be found. It can contain at most two electrons.**
    - **Any** s sublevel has 1 orbital
        - 1s has one orbital
        - 2s has one orbital
        - 3s has one orbital
        - 10s has one orbital, etc.
        - An s sublevel is spherical in shape



    - Any s sublevel has one orbital, which can hold 2 electrons at most. Therefore, any s sublevel can hold at most 2 electrons.
- Any p sublevel has 3 orbitals.
    - 2p, 3p, 4p, etc. all have three orbitals each.
    - p orbitals look like dumbbells, and are right angles to one another in 3-dimensional space. They can be thought of as being oriented on three axes (x, y, and z).
    - They are called the p_x, p_y, and p_z orbitals.
    - Because any orbital can contain at most 2 electrons, and because any p sublevel has three orbitals, **any p sublevel can contain at most 6 electrons.**



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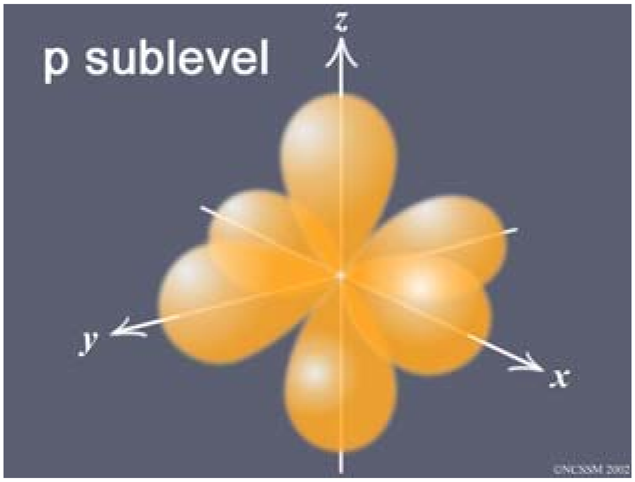
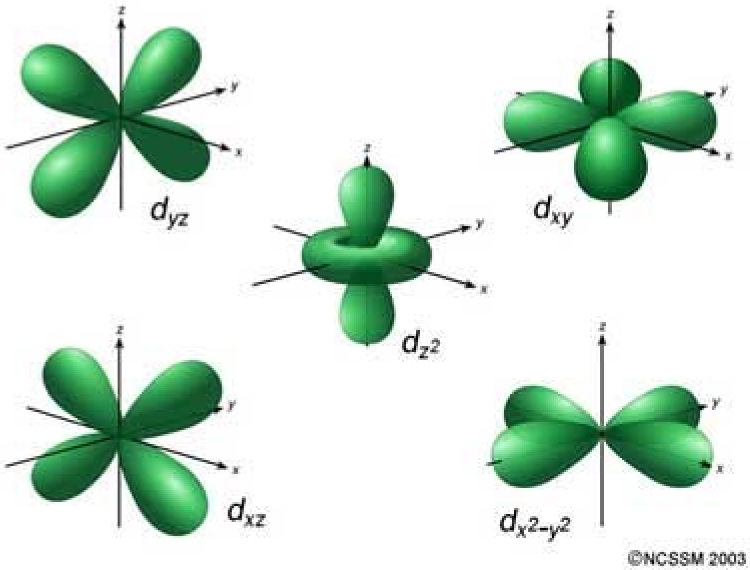
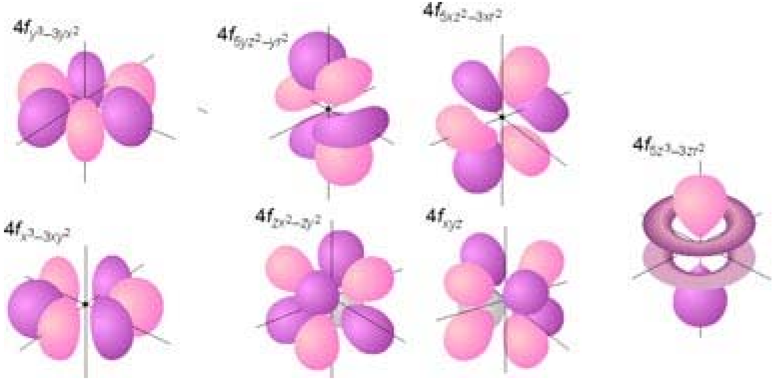


- Any d sublevel has 5 orbitals.
    - The 3d sublevel has 5 orbitals. So does the 4d and so on.
    - That means that any d sublevel can hold a maximum of 10 electrons.
    - The shapes and orientations of the d orbitals are complex.



- Any f sublevel has 7 orbitals.
  - The 4d sublevel has 7 orbitals. So does the 5d and so on.
  - That means that any f sublevel can hold a maximum of 14 electrons.
  - The shapes and orientations of the f orbitals are complex.



- To summarize,

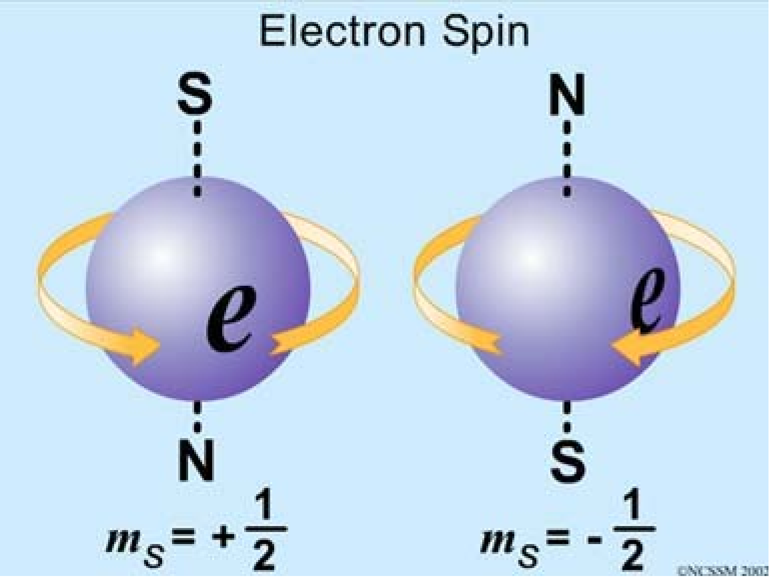
Sublevel type	# of orbitals	Max # of electrons in sublevel
s	1	2
p	3	6
d	5	10
f	7	14
g, but not important to us	9, but not important to us	18, for the sake of demonstration
Who cares	11, but who cares	22, but don’t lose sleep

Energy level (=n)	Sublevels present	# of orbitals ($=n^2$)	Max # of electrons ($=2n^2$)
1	1 (1s)	**1**	**2**

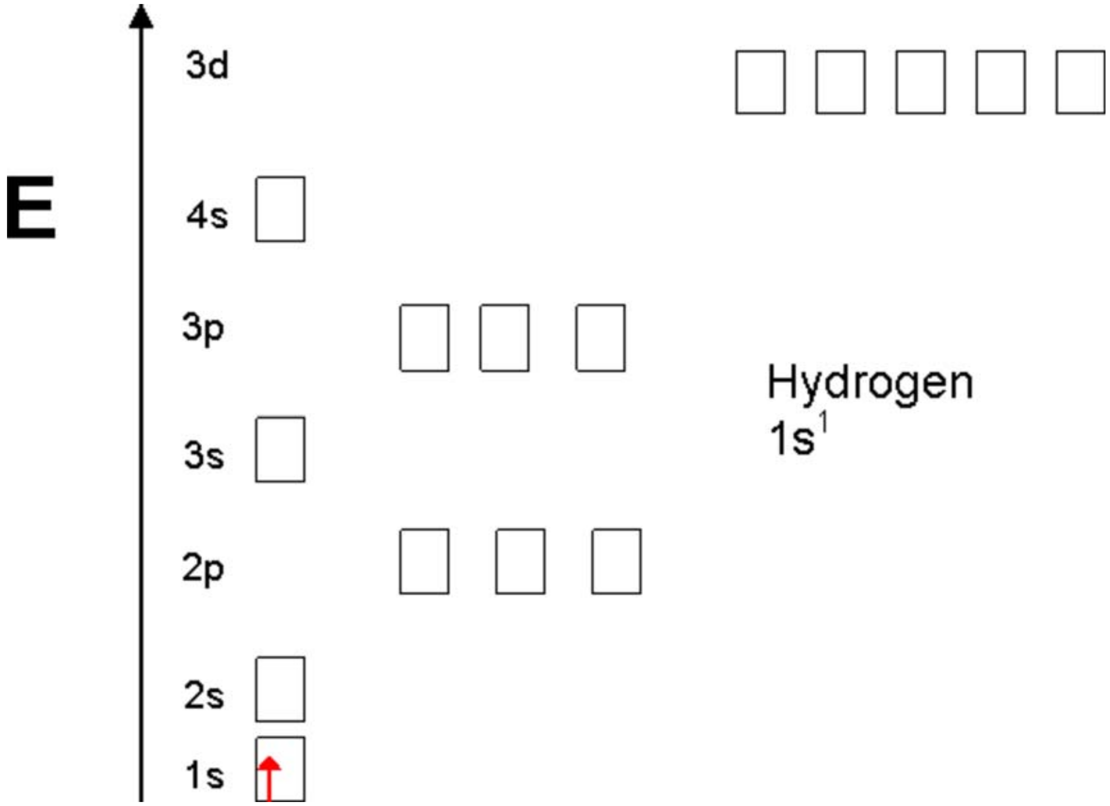
2	2 (2s, 2p)	1+3 = **4**	2 + 6 = **8**
3	3 (3s, 3p, 3d)	1+3+5 = **9**	2 + 6 + 10 = **18**
4	4 (4s, 4p, 4d, 4f)	1+3+5+7 = **16**	2 + 6 + 10 + 14 = **32**
5	5 (5s, 5p, 5d, 5f, 5g)	1+3+5+7+9= **25**	2 + 6 + 10 + 14 + 18 = **50**
6	6, etc.	1+3+5+7+9+11= **36**	2 + 6 + 10 + 14 + 18 +22= **72**
7	7, etc.	1+3+5+7+9+11+13= **49**	2 + 6 + 10 + 14 + 18 +22 + 26 = **98**

- Don’t memorize any of this info, except possibly the n^2 and $2n^2$ shortcuts, because this info is all summarized on the periodic table!
- Demonstration: determining the available sublevels in the first 4 energy levels by looking at the periodic table.
- Demonstration: determining the number of electrons per each orbital type by looking at the periodic table.
- Demonstration: determining the number of orbitals per each sub level type by looking at the periodic table.

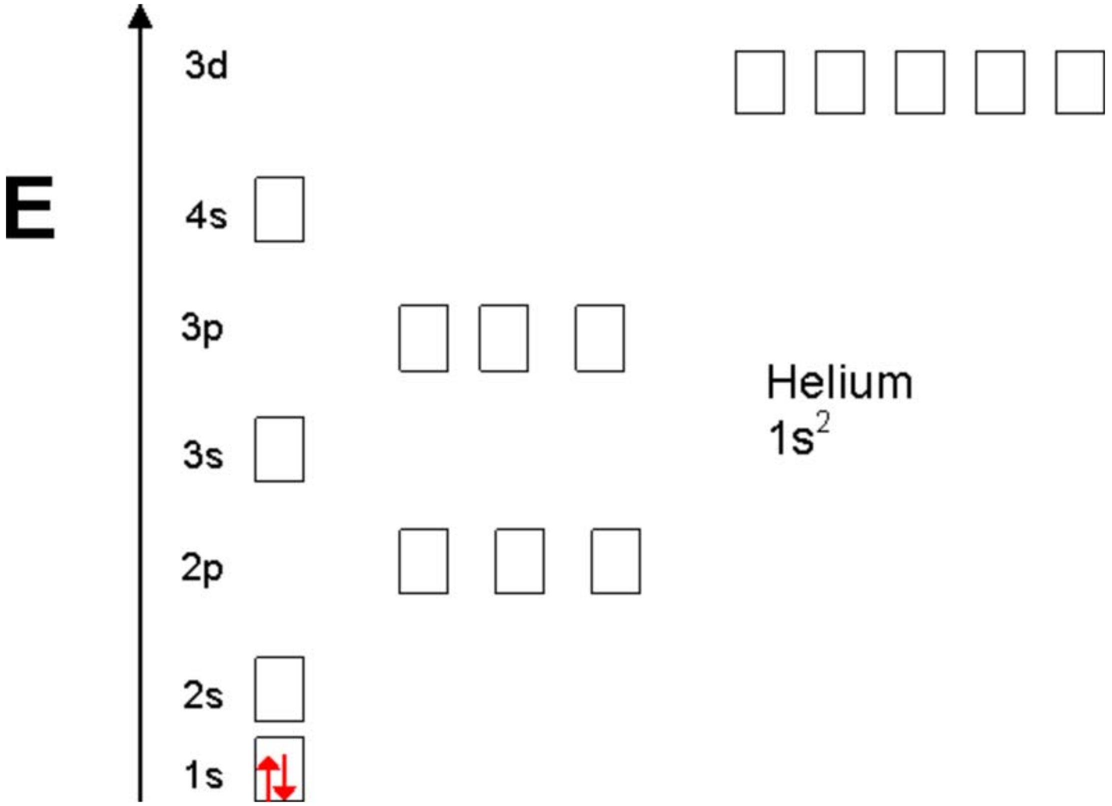
- Time for a definition of “orbital”
    - An orbital is a region in space where an electron is likely to be found
    - Orbitals don’t need to have any electrons, but can contain at most 2 electrons. That is, an orbital can have 0, 1, or 2 electrons
    - To repeat and emphasize, **ANY orbital - an s orbital, a p orbital, a d orbital ,etc. - can contain at most 2 electrons.**
    - If an orbital has two electrons, the spins must be opposite
        - “Spin” is a property of electrons
        - In reality, the electrons do not spin, but it is convenient to describe this property as spin.
        - The values of m_s can be either + 1/2 or -1/2.
        - We will call these “spin up” () and “spin down” (¯).



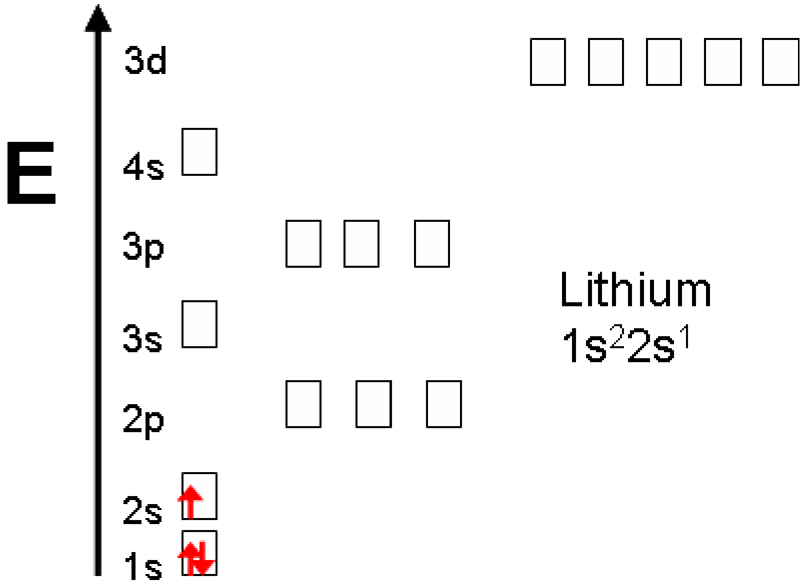
- Now let's look at a simplified energy diagram of the atomic orbitals, arranging them from the first to be filled to the last to be filled.
- Note: the orbitals within a given energy level all have the same energy. However, for reasons we will not cover in this course, the orbitals are filled in the order shown below.
- Note: the 4s and 3d sublevels "overlap"; that is, the 4th energy level (4s) starts filling before the 3rd energy level is full of electrons.
- Note: no atom is actually formed in this way. However, in order to understand why the electrons are arranged as they are in an atom, it is often helpful to pretend that the electrons are added to the electron cloud one at a time.
- This diagram is called an Aufbau diagram.
- Let's start with the easiest atom: hydrogen. Hydrogen is atomic #1, thus a neutral atom of H has one electron.



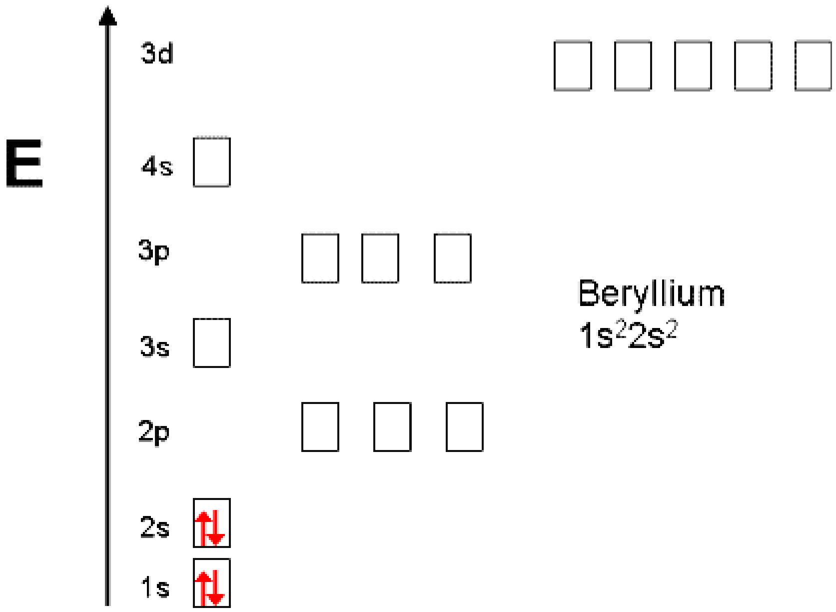
- Because H has one electron in the 1s sublevel, its electron configuration is $1s^1$. The Aufbau diagram for an element is a picture (shown above). The electron configuration is a string of numbers and letters.
- The rules for filling the Aufbau diagram are as follows:
    - The Aufbau principle:
        - Electrons enter orbitals of lowest energy first.
    - Pauli's exclusion principle:
        - Orbitals may contain at most two electrons, and if there are two electrons, they must have opposite spins. ("No two electrons can have the same four quantum numbers.")
    - Hund's Rule
        - When electrons enter orbitals of equal energy, orbitals must first be singly occupied with spins parallel before they are doubly occupied.
- Now let's look at the Aufbau diagram and electron configuration for helium, He, Z=2.



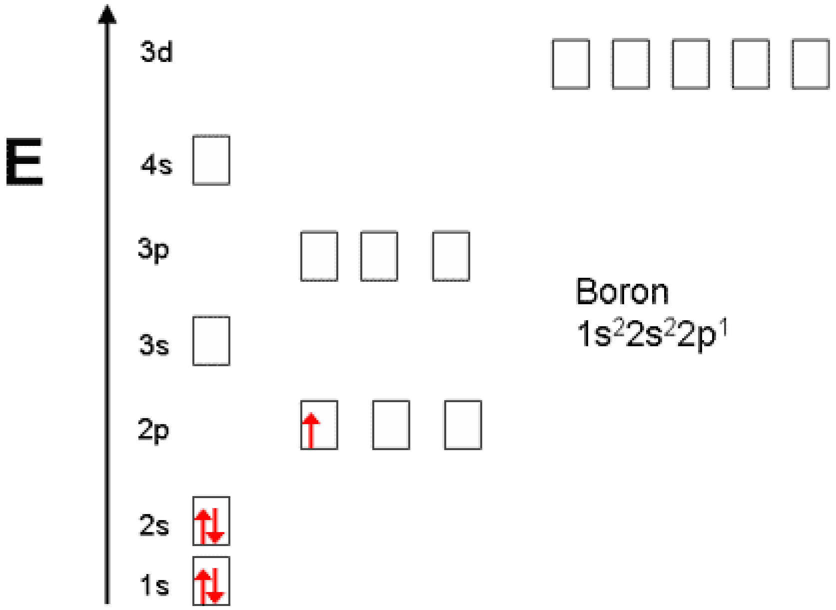
- Here is the diagram of lithium:



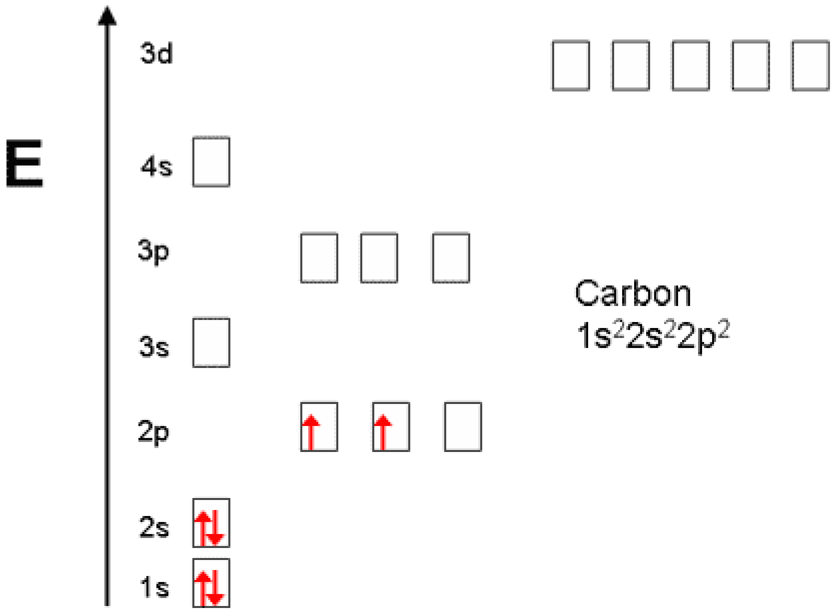
- Here is the diagram for Be, Z=4:



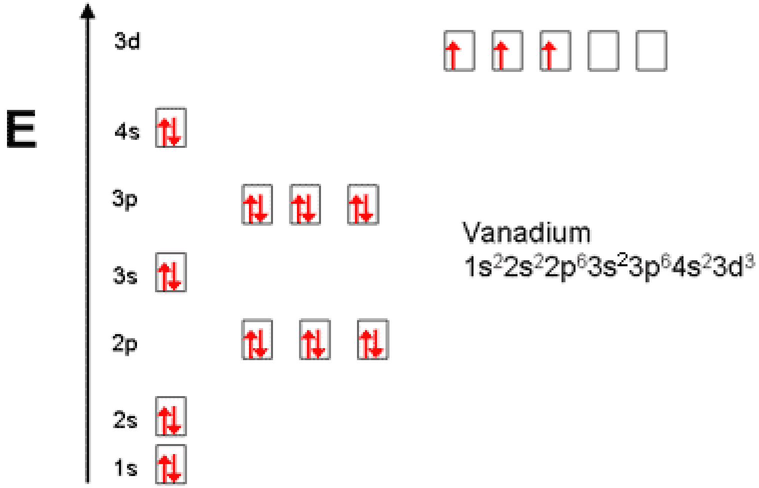
- Here is the diagram for B (boron), Z=5:



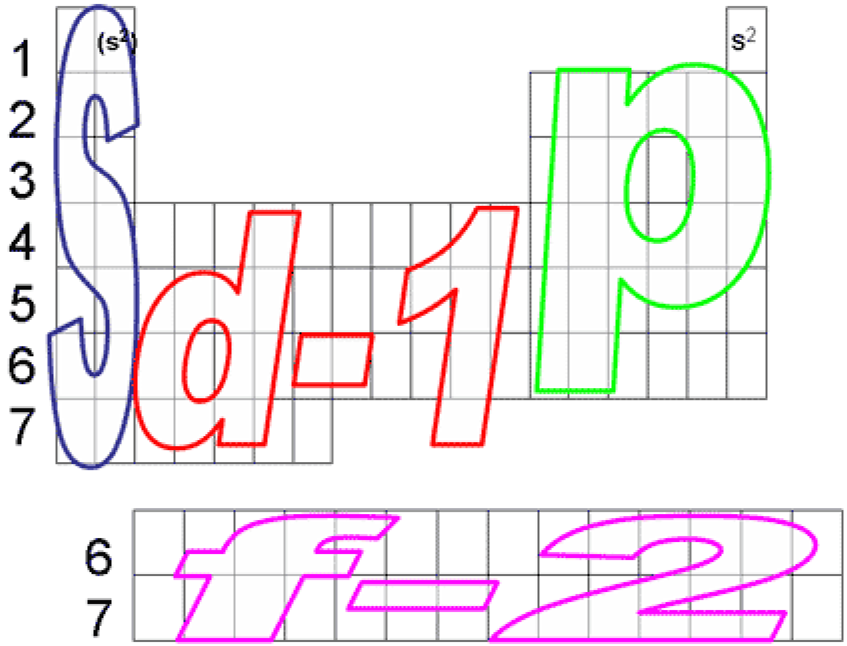
- Now we need to invoke Hund's rule to figure out where the next electron will go. Here is the electron configuration and Aufbau diagram for carbon (Z=6):



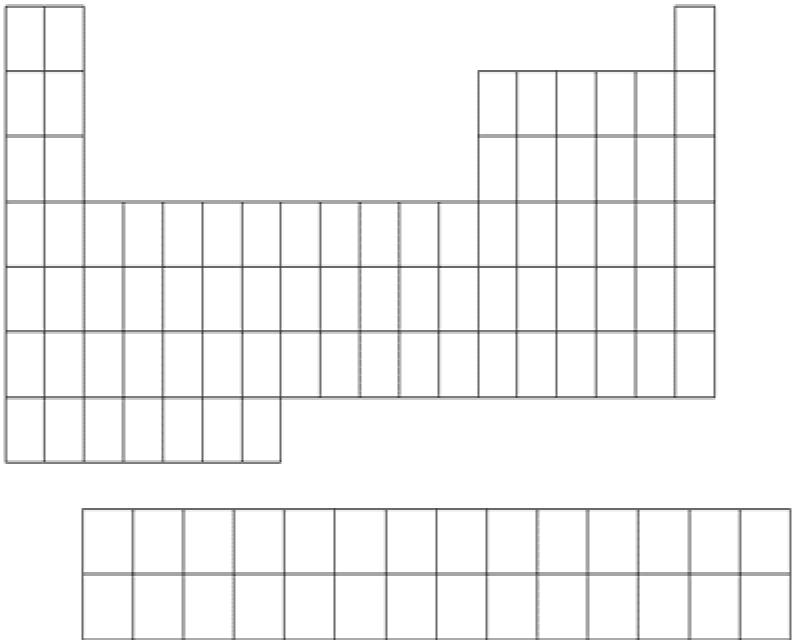
- Here is vanadium (V, Z=23).



- How to use the periodic table to determine the electron configuration: the order of orbital filling is summarized by the guide shown below:



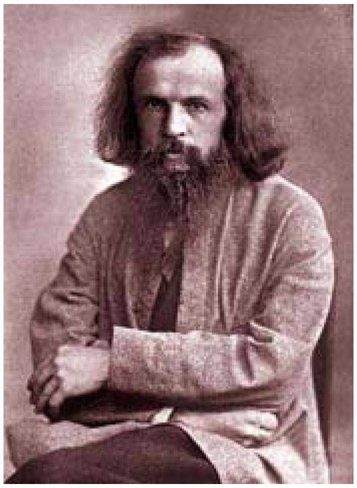
- See if you can determine the electron configuration of the elements from the above guide. There are exceptions to this rule, but the exceptions are unrelated to the order of filling of the orbitals.



- Noble gas abbreviation:
    - The electron configuration for an element can be abbreviated by putting the symbol of the last noble gas before that element occurs in the periodic table in brackets, and then adding the “remainder” of the configuration notation after that.
    - Use the “Last Noble Gas That You Passed”
    - Sodium is $1s^22s^22p^63s^1$, or simply $[Ne]3s^1$.
    - Calcium is $1s^22s^22p^63s^23p^64s^2$, or simply $[Ar]4s^2$.
    - Chlorine is $1s^22s^22p^63s^23p^5$, or simply $[Ne]\ 3s^23p^5$.
    - Radium is $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^66s^24f^{14}5d^{10}6p^67s^2$, or simply $[Rn]7s^2$.
- Valence Electrons:
    - The elctrons that are involved in bonding.

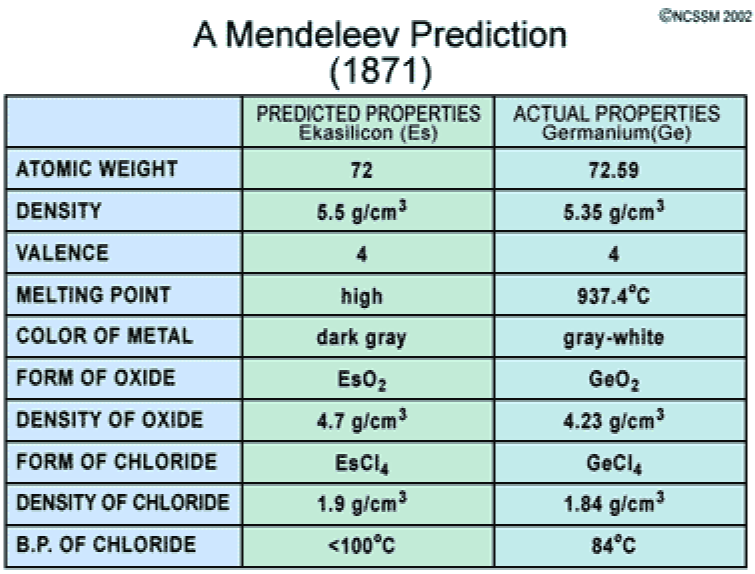
    - Generally, they are the s and p electrons in the highest occupied energy level.
- Exceptional configurations:
    - Some electron configurations violate the Aufbau principle
    - Many inner transition metals (a.k.a. rare earth elements, a.k.a. f-block elements) do this.
    - You will be responsible for examples found in the transition elements (a.k.a., the “d” block).
    - Common examples that you should know: s^2d^4 and s^2d^9 configurations do not occur.
    - They instead occur as ~~s^2d^4 and s^2d^9~~ s^1d^5 and s^1d^{10}.
    - An s electron is promoted to the nearby d orbital whenever the d orbital is one electron away from being filled (10 electrons) or half-filled (5 electrons).
- Examples:
    - Copper (Cu) does not end in “$4s^23d^9$”. Instead, $[Ar]4s^13d^{10}$.
    - Silver (Ag) does not end in “$5s^24d^9$”. Instead, $[Kr]5s^14d^{10}$.
    - Chromium (Cr) does not end in “$4s^23d^4$”. Instead, $[Ar]4s^13d^5$.
    - Molybdenum (Mo) does not end in “$5s^24d^4$”. Instead, $[Kr]5s^14d^5$.

- Periodic table: table showing all of the elements in increasing atomic number.
- Development of the periodic table
    - **Our current periodic table is not the first periodic table.**
    - What does the word “periodic” mean? “periodical”? “periodicity”?



Dmitiri Mendeleev, 1834-1907

- The first periodic table was devised by a Russian chemist, Dmitri Mendeleev. **Mendeleev arranged his periodic table by increasing atomic mass.**
- Mendeleev was writing a chemistry textbook, and wanted to organize the information that was known about all of the known chemical elements. At the time, scientists had isolated and identified about 70 elements.
- The story goes that he wrote the symbol of each element on a card, and then wrote the chemical and physical data about each chemical on the back of each card.
- As he arranged the cards in order of increasing atomic mass, he started to notice certain recurring patterns in the elements’ properties.
- These properties re-occurred at regular intervals. For instance (this is a simplification), when he arranged the elements in order of increasing atomic mass, he noticed that every eight elements, and then every eighteen elements, there was an explosive metal (Na, K, Rb, etc.)
- When you lined up a set of such elements into a “family” according to shared characteristics, other sets of elements with shared characteristics also fell into line.
- Mendeleev was able to see this pattern and use these patterns to predict the existence and properties of as-yet undiscovered elements.
- Mendeleev was able to win respect for his bold predictions when these elements were discovered and studied. Examples: eka-silicon (germanium), eka-aluminum (gallium), and eka-boron (scandium).



A Mendeleev Prediction
(1871)

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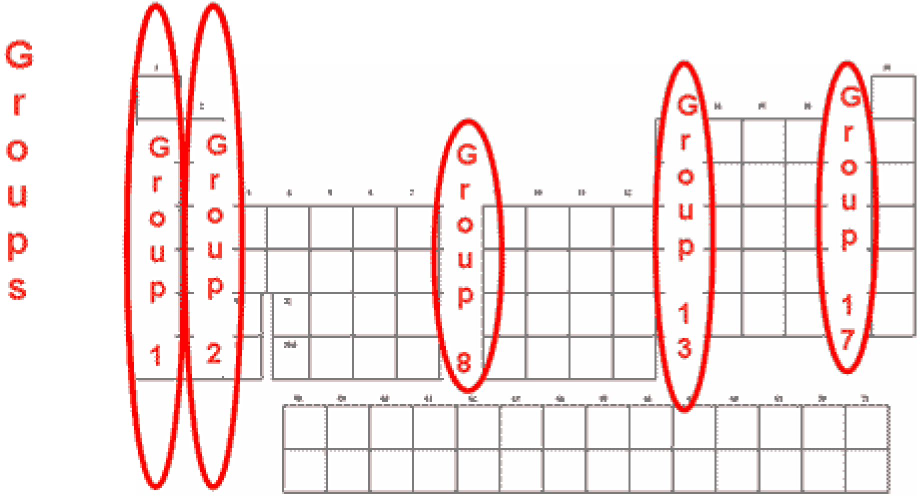
	PREDICTED PROPERTIES Ekasilicon (Es)	ACTUAL PROPERTIES Germanium(Ge)
ATOMIC WEIGHT	72	72.59
DENSITY	5.5 g/cm^3	5.35 g/cm^3
VALENCE	4	4
MELTING POINT	high	937.4°C
COLOR OF METAL	dark gray	gray-white
FORM OF OXIDE	EsO_2	GeO_2
DENSITY OF OXIDE	4.7 g/cm^3	4.23 g/cm^3
FORM OF CHLORIDE	$EsCl_4$	$GeCl_4$
DENSITY OF CHLORIDE	1.9 g/cm^3	1.84 g/cm^3
B.P. OF CHLORIDE	<100°C	84°C

  - He was able to predict the properties of these elements by interpolating the patterns of the surrounding elements on his periodic table. He also stated that several elements had incorrectly stated atomic masses, since they did not fit into the correct spot on his table. He was mostly right on this point, but not always (see below).
  - Fun fact: he wrote his doctoral thesis, essentially, about vodka ("On Composing Alcohol with Water").
- **There are problems with arranging the periodic table by increasing atomic mass.** Can you find examples on our periodic table where the elements are not in order of lightest to heaviest?
- It is obvious by looking at iodine’s chemical properties that it belong in its group, Group 7A (a.k.a. Group 17). For instance, it reacts with Na in a 1:1 ratio, making NaI (just like NaCl, NaBr, NaF, etc.).
- However, if you put iodine in the correct group, then it is out of order with respect to mass. If you put Iodine in the “correct” spot according to its mass, then it winds up in the wrong family.
- Henry Moseley helped to develop the modern periodic table.
- Moseley was able to determine the “atomic charge” of an element, what we would today call the number of protons. This atomic charge is now called the atomic number.
- **Henry Moseley is responsible for today’s arrangement of the periodic table. Today the periodic table is arranged by increasing atomic number.**

- If the elements are arranged by increasing atomic #, then there are no “mistakes” in how the elements are organized into the proper groups.
- The periodic table
    - **Groups, a.k.a. families** – these are the vertical columns of elements. Elements in the same families have similar chemical and physical properties.



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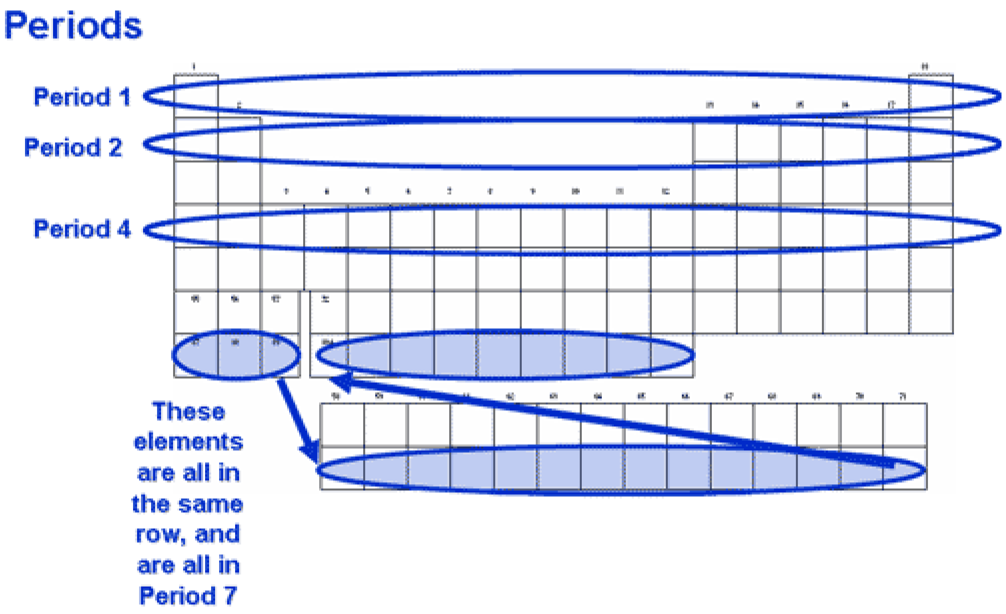
    - **Periods –** these are the horizontal rows of the periodic table. Elements in the same period have very different properties, for the most part.



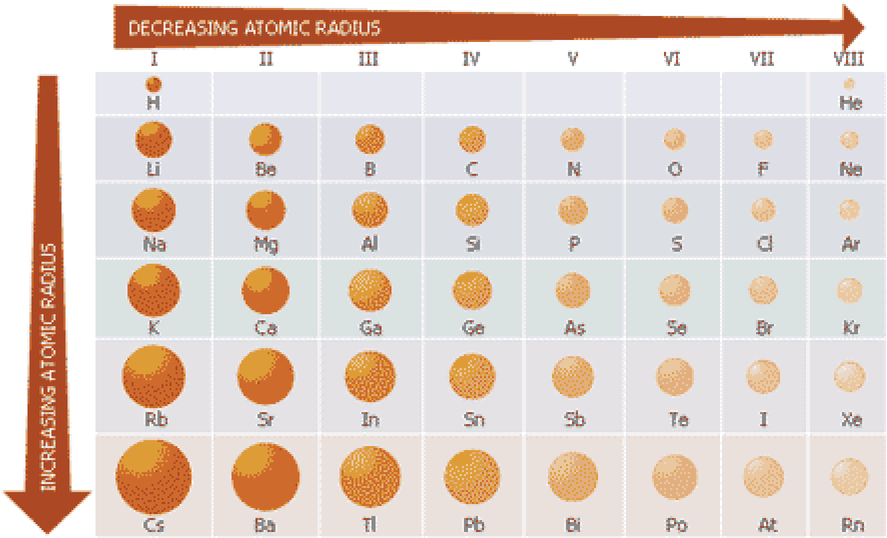
- The periodic table has three main sections: The metals (left side of staircase), the nonmetals (right side of staircase), and the metalloids (touching the staircase with a whole side of their square).
- Exceptions: Hydrogen is a nonmetal. Aluminum is a metal.
- **Metals** (p.103)
  - Good conductors of heat & electricity
  - Malleable
  - Ductile
  - Lustrous
- **Nonmetals**
  - Poor conductors (**insulators**)
  - Brittle
  - Dull
- **Metalloids**
  - Properties that are somewhere in between those of metals and nonmetals
  - Example: silicon (Si). It is shiny but brittle. It is a **semiconductor**, which is a material that is neither a good conductor nor good insulator. It can control the flow of electrons (electricity).
- Numbering system of the groups: we will use the easy, new system: 1 through 18.
- Group names that you should know:
  - Alkali metals (group1)
  - Alkaline earth metals (group 2)
  - Chalcogens (group 16)
  - Halogens (group 17)
  - Noble gases, a.k.a. inert gases (group 18)
  - Transition metals (d-block elements)
  - Inner transition metals (f-block elements)
- Elements found alone in nature (uncombined):
  - Noble gases
    - Unreactive elements, always found alone in nature
  - Gold, Silver, Platinum
    - Relatively unreactive metals, often found chemically pure (or nearly so) in

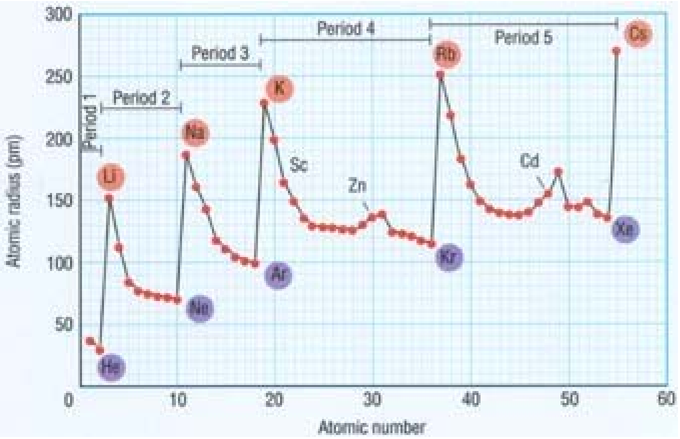
nature

- Most other elements are found combined with other elements as compounds.
- Some elements combine with themselves to form diatomic molecules:
  - H_2
  - N_2
  - O_2
  - F_2
  - Cl_2
  - Br_2
  - I_2

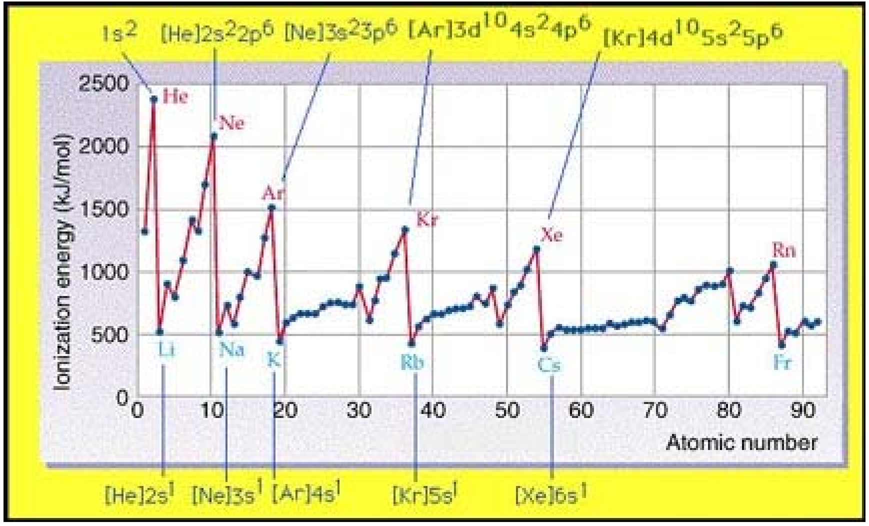
Periodic Trends

- **The Periodic Law:** there is a periodic repetition of the physical and chemical properties of the elements when they are arranged in order of increasing atomic number.
- There are a number of atomic properties that vary in a regular way across the periods and down the groups.
- Atomic radius
  - Atomic radius is the distance from the center of the atom to the outer edge of the atom
  - It is a way of measuring the size (in distance) of atoms
  - The periodic and group trends are
    - L to R, the atoms decrease in size
    - T to B, the atoms increase in size
  - Reasons for the trend:
    - As you go down a group, the number of “shells” or energy levels of electrons between the positive nucleus and the negative nucleus increases.
    - This decreases the attraction between the nucleus and the outermost electrons.
    - The effect is that the electrons are less tightly held, and therefore are able to get further away. This is called the **shielding effect**.
    - As the atomic number across a period increases (L to R), the number of electrons (in the outermost energy level) and the number of protons (in the nucleus) both increase.
    - Without the shielding effect of adding more “layers” of electrons between the nucleus and these additional electrons, the net effect is to draw the outermost electrons closer to the nucleus.
    - “more protons + more electrons = stronger attraction”: . . . within a given period.

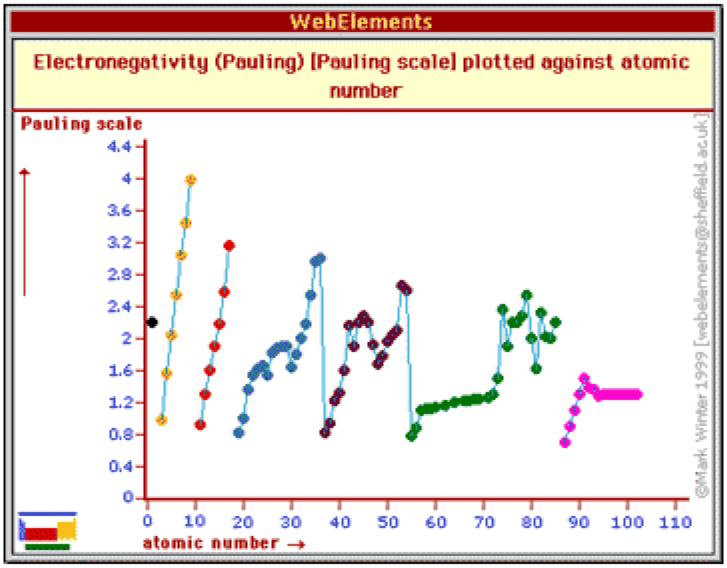




- Ionization energy
  - Ionization energy is the energy required to remove an electron from an atom
  - Metals give up electrons more easily than nonmetals; metals will attain a full octet by losing relatively few electrons.
  - Nonmetals will attain a full octet more easily by gaining electrons than by losing them.
  - Thus, the trends are
    - L to R ionization energy increases
    - T to B ionization energy decreases



- Electronegativity
  - Electronegativity is the tendency of an atom to attract electrons toward itself when bonded to another atom
  - Metals have low electronegativities
  - Nonmetals have high electronegativities
  - Thus, the trends are
    - L to R electronegativity increases
    - T to B electronegativity decreases


WebElements
Electronegativity (Pauling) [Pauling scale] plotted against atomic number
Pauling scale
4.4
4
3.6
3.2
2.8
2.4
2
1.6
1.2
0.8
0.4
0
0
10
20
30
40
50
60
70
80
90
100
110
atomic number →
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