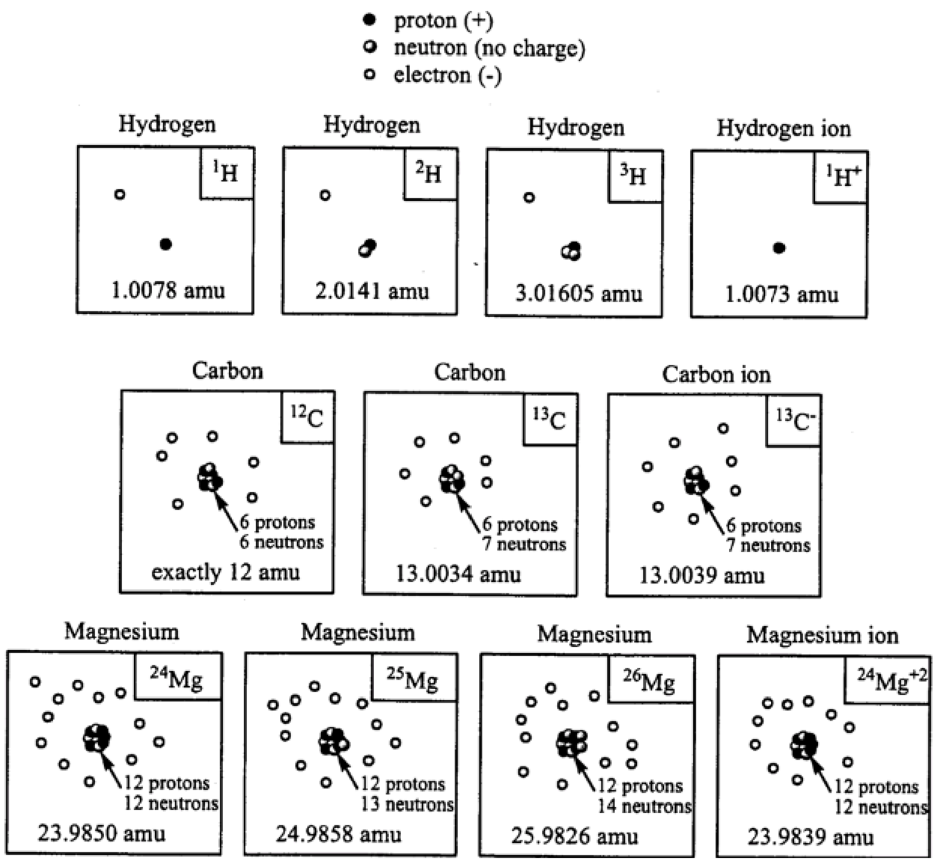
ATOMS, ISOTOPES, AND IONS

Particle	Mass (amu)	Charge
Proton	1.0073	+1
Neutron	1.0087	0
Electron	0.0005	-1

amu = atomic mass unit

Model 1: Schematic diagrams for various atoms



^1H, ^2H, and ^3H are **isotopes** of hydrogen. ^{12}C and ^{13}C are **isotopes** of carbon.
^{12}C and ^{13}C may also be written as "carbon-12" and "carbon-13"
The nucleus of an atom is extremely small and dense.
It contains the protons and the neutrons (if any).

Critical Thinking Questions

1. What are the three particles that comprise an atom?

2. Which particles contribute most of the mass to the atom, and where are these particles located?

3. Which particles contribute most to the volume/size of the atom and where are they located?

4. Look at the Model 1 for carbon. What do all three carbon atoms (and ions) have in common?

5. What do all four hydrogen atoms and ions have in common?

6. What do all magnesium atoms and ions have in common?

7. Look at a periodic table. What is similar to your answer in CTQ 4-6 with carbon, hydrogen, and magnesium on the periodic table? This is called the **atomic number**. What is the significance of the atomic number and what does it tell you about the element?

8. How many protons are in all chlorine atoms? ______ Do you think chlorine atoms exists with 18 protons? _____ Why or why not?

9. What is the significance of the superscripted + or – symbols to the right of the symbols in terms of the number of protons and electrons

   a. for the hydrogen ion (H^+)?

   b. for the carbon ion (C^-)?

   c. Write one complete sentence to define the term ION.

10. In a box in the top right corner of each schematic in Model 1 is the element symbol and the **mass number** for the atom (superscript on the left side of the element symbol). How is the mass number determined?

11. What is the mass number for the following atoms:

a. ^{37}Cl ________ b. ^{238}U________ c. carbon-12 ________ d. carbon-13 ________

12. What structural element do all isotopes of an element have in common? How are their structures different?

13. Considering what you know about isotopes, do all atoms of an element weigh the same? Explain.

Problems

1. Complete the following table. The first one has been done for you.

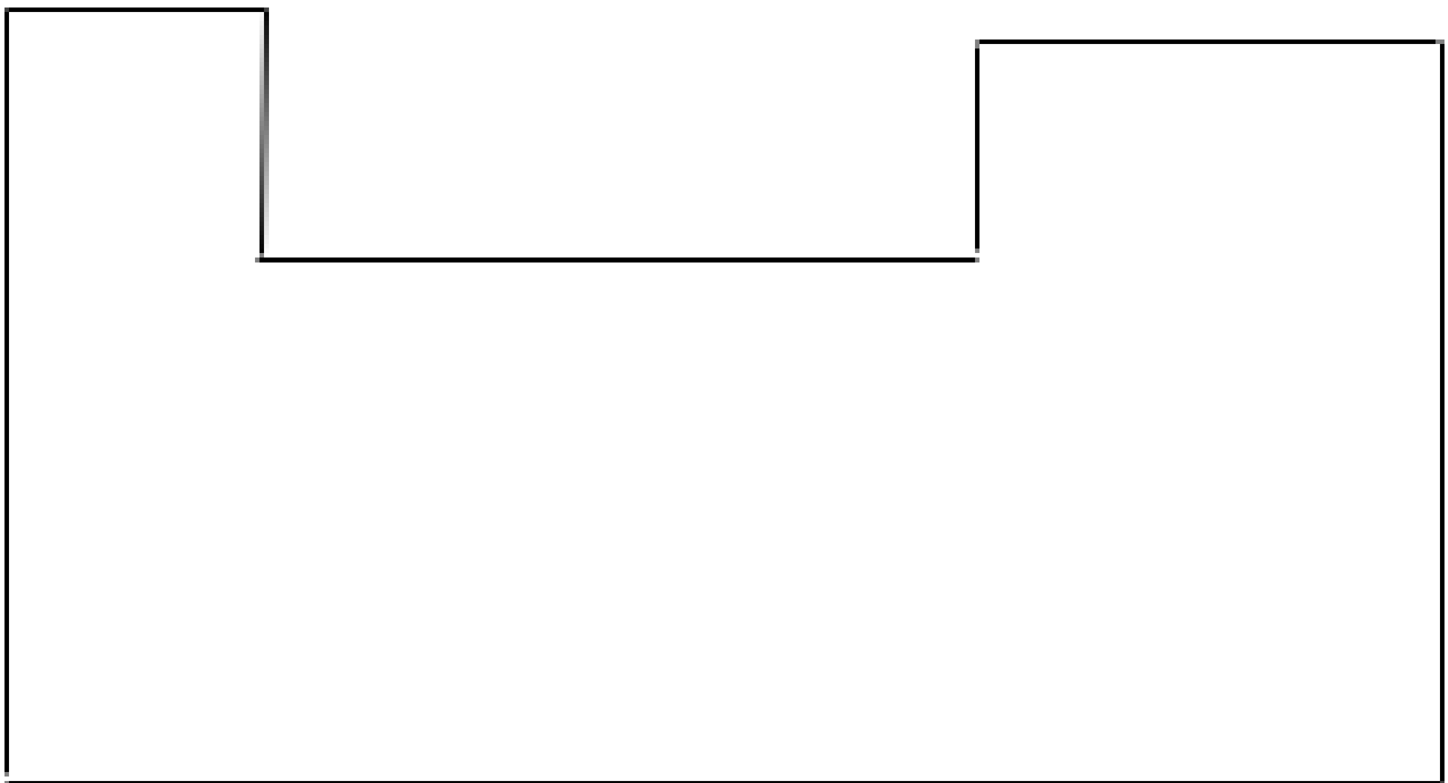
Symbol	Atomic #	Mass #	# Protons	# Neutrons	# Electrons	Overall Charge
^{16}O						
^{18}O						
				146	92	
	17			18	18	
		64			29	0
		40			18	2+
	5			5	4	
$^{32}S^{2-}$						

What is the difference between an ATOM and an ELEMENT?

THE PERIODIC TABLE

Information:

The Periodic Table lists all the known elements in order of their atomic number and in columns that depend on similarities in their chemical and physical properties. The Periodic Table is a useful tool for both students and professionals to identify the properties of the elements and understand the properties of molecules.



Model 1: Where everything is located on the Periodic Table

Groups: columns on the Periodic Table (numbered 1-18)

Elements with similar properties are placed in groups

Groups 1, 2, and 13-18 are called the Main Group Elements
Group 1 – Alkali Metals – forms +1 charge
Group 2 – Alkaline Earth Metals – forms +2 charge
Group 14 – forms +4 or –4 charge
Group 15 – forms –3 charge
Group 16 – forms –2 charge
Group 17 – Halogens – forms –1 charge
Group 18 – Noble Gas – doesn't form a charge
Group 3 – 12 – Transition Metals – variable charges

Periods: rows on the Periodic Table (numbered 1-7)

LOCATION: Elements on the periodic table are either classified as metals, non-metals or metalloids/semimetals.

Metals: Left of the Periodic Table

Ability to be pulled into a wire (ductility)
Good conductors of heat and electricity
Shiny and found as a solid at room temperature (except Mercury, Hg)

Nonmetals: Right of the Periodic Table

Bad conductors of heat and electricity (good insulator)
Can be a solid, liquid, or gas

Semiconductors/Semimetals/Metalloids: Staircase line dividing the nonmetals and metals

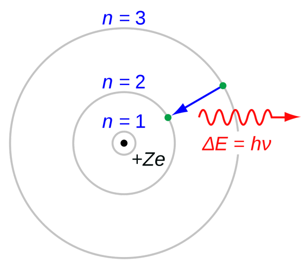
Have some of the same properties as metals and nonmetals
Dull in color
Usually solid at room temperature
Can conduct heat and electricity, but not as well as metals

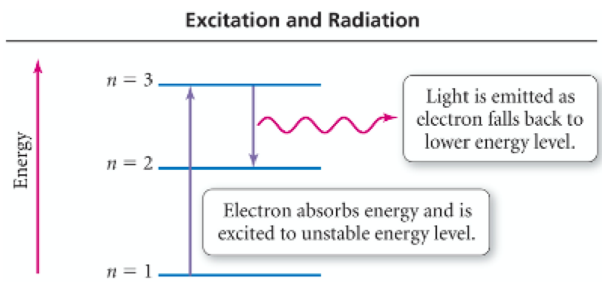
THE DESCRIPTION OF ELECTRONS IN ATOMS

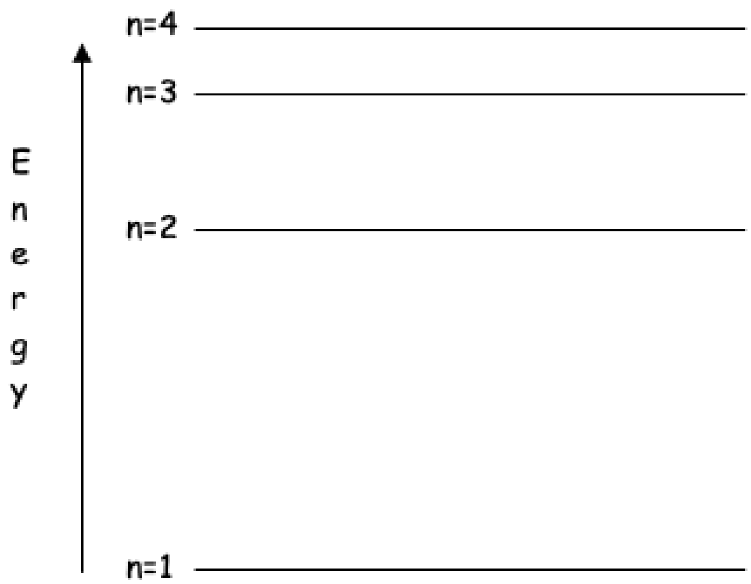
To understand electron configuration, we must understand the importance of electrons and its components.The Bohr Model of an Atom offers a backbone in understanding electrons and their atomic structure. Breaking down the structure of an atom, we can further locate electrons in an atom.
There are two methods of notating electron configuration: orbital notation and atomic notation.

The Bohr Model of an Atom
The Bohr Model of an Atom defines an atom as a small, positively charged nucleus surrounded by negatively charged electrons that orbit the nucleus on shells. Shells are energy levels of an atom.
The Bohr Model focuses on the atomic structure of electrons of an atom. See Atoms, Isotopes and Ions Activity

Example: fireworks







The Bohr Model is a planetary model in which the negatively charged electrons orbit a small, positively charged nucleus similar to the planets orbiting the sun (except that the orbits are not planar). The gravitational force of the solar system is mathematically akin to the Coulomb (electrical) force between the positively charged nucleus and the negatively charged electrons.

Main Points of the Bohr Model

Electrons orbit the nucleus in orbits that have a set size and energy.

The energy of the orbit is related to its size. The lowest energy is found in the smallest orbit.

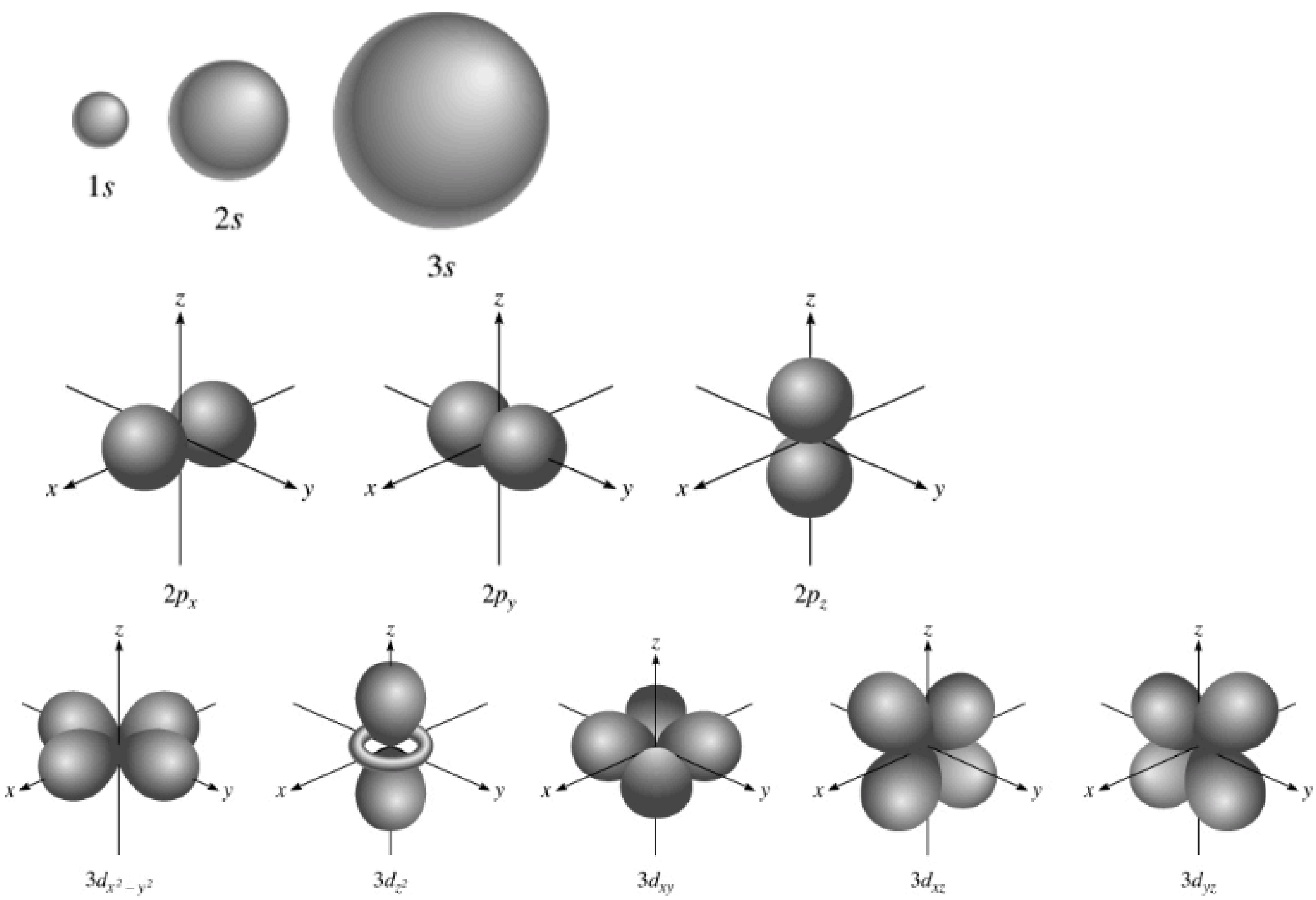
Radiation is absorbed or emitted when an electron moves from one orbit to another.

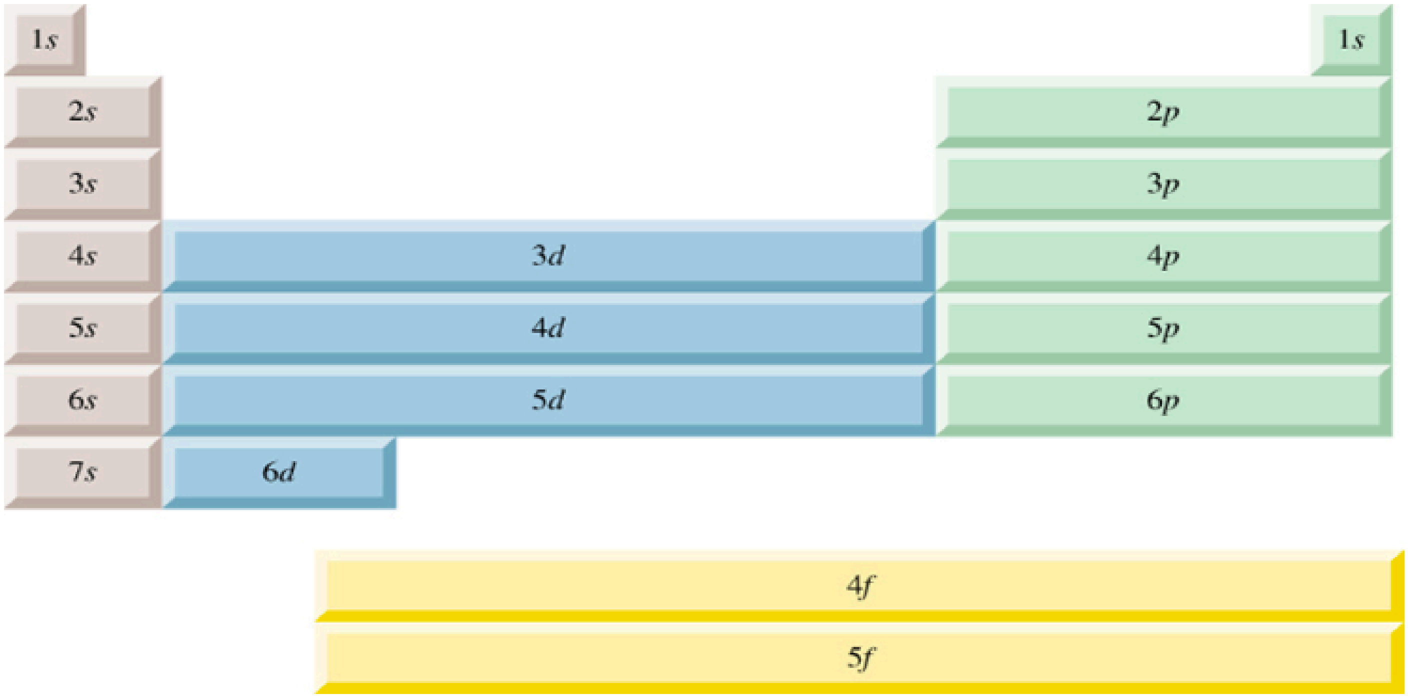
Problems With the Bohr Model

The **Bohr model** fails to describe the absorption/emission spectra of atoms that are more complex than hydrogen. It violates the Heisenberg Uncertainty Principle because it considers electrons to have both a known radius and orbit.
It makes poor predictions regarding the spectra of larger atoms.
It does not predict the relative intensities of spectral lines.

The **quantum mechanical model** of the atom uses complex shapes of orbitals (sometimes called electron clouds), volumes of space in which there is likely to be an electron. So, this **model** is based on probability rather than certainty.

Electrons are located in regions of space that are called ORBITALS. *The atomic orbital is the region of space where an electron can be found.* An orbital can be represented by drawing a boundary surface to identify that the electron has a 90% probability of being within that surface.





Shorthand Notation

S $1s^22s^22p^63s^23p^4$ "e^- configuration" the noble gas right before it: Ne $1s^22s^22p^6$

$[Ne]3s^23p^4$ "shorthand notation"

As $1s^22s^22p^63s^23p^64s^23d^{10}4p^3$

$[Ar]\ 4s^23d^{10}4p^3$

Information

Core electrons – e^- up to the noble gas and including d electrons – inner shell e^-

(in As - $1s^22s^22p^63s^23p^63d^{10}$, in S - $1s^22s^22p^6$)

Valence electrons – e^- that does bonding, outer shell e^-

(in As - $4s^24p^3$, therefore 5 VE, in S – $3s^2\ 3p^4$, therefore 6 VE)

<u>e^- configuration for ion</u>

Ca^{2+} $1s^22s^22p^63s^23p^6$
Cl^- $1s^22s^22p^63s^23p^6$
K^+ $1s^22s^22p^63s^23p^6$
Ar $1s^22s^22p^63s^23p^6$

Note: same e^- configuration, called isoelectronic series

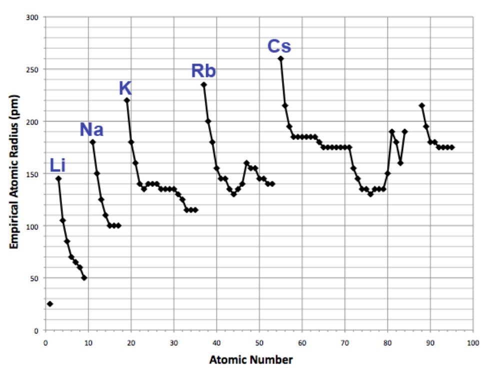
PERIODIC TRENDS

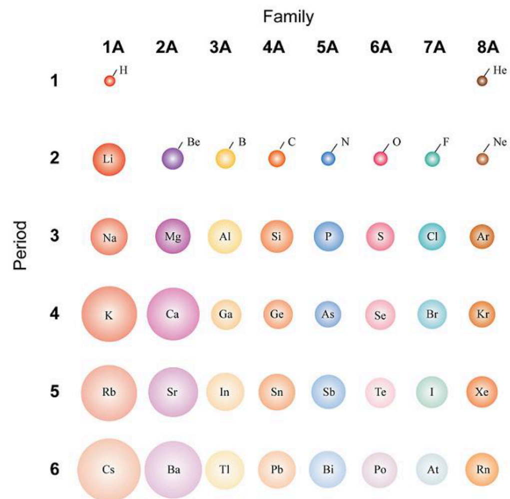
Why?

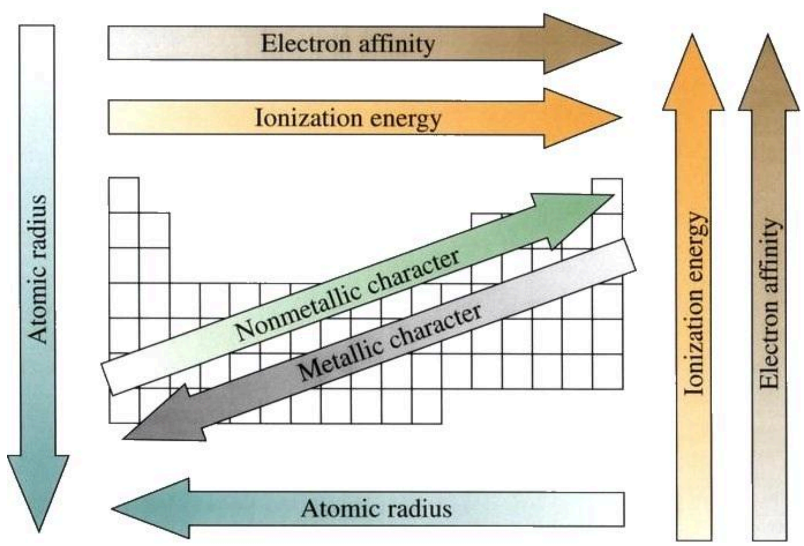
Properties such as the size of an atom (atomic radius), the energy to remove an electron from an atom (ionization energy), the energy change when an electron is added to an atom (electron affinity), and the ability of an atom to attract electrons to itself, can be understood in terms of electron configuration and the competition between electron-nucleus attraction and the electron-electron repulsion.

Model 1 Trends

I. Atomic Radii







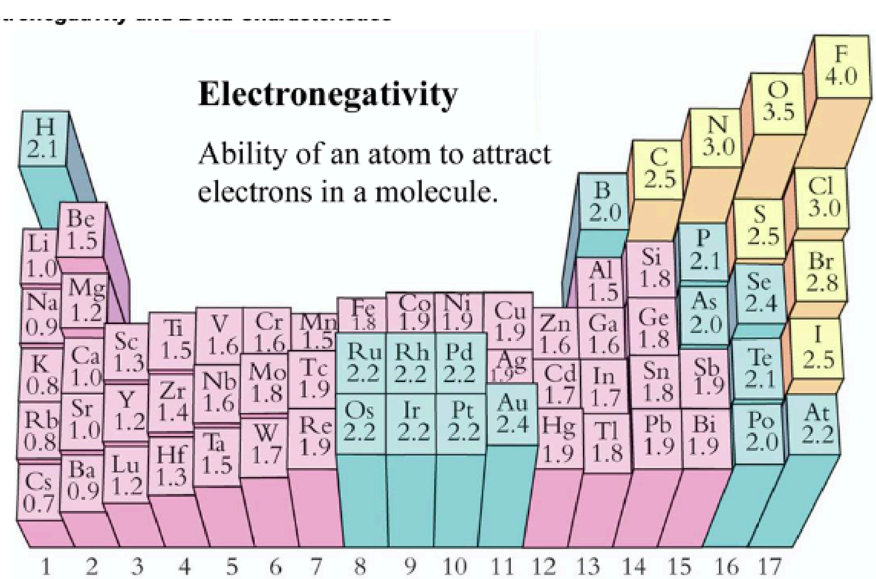
Effective Nuclear charge (Z_{eff}) - The larger an atomic radius, presumably the more shells of electrons the atom has. Therefore, there are also more electrons on the inner shells that shield the electrons in the outer shells from the full electrostatic charge of the nucleus, so the effective charge (Zeff) on them is reduced. If there is a high effective charge on an electron, it is not being shielded by many (or possibly any) electrons that are of lower shells/closer to the nucleus, so therefore the pull on the electron is very strong and it stays closer to the nucleus, decreasing atomic radius.

Ions

Ionization energy amount of energy needed to take an electron away from a gaseous atom

Electron affinity   amount of energy released when an atoms takes an electron in the gaseous state.

Electronegativity



LEWIS MODEL OF ELECTRONIC STRUCTURE

Information

Molecules exist because they are more stable than the separate atoms. By "more stable", we mean that they have lower energy. G.N. Lewis recognized from the very low chemical reactivity of the noble gases that a configuration with $8e^-$ in a shell produces a very stable situation. He therefore proposed that molecules form so that atoms can transfer or share electrons and produce this very stable octet structure.

In the Lewis structure diagrams, dots represent electrons, a line between the atom represents a single covalent bond formed by a pair of electrons, other dots represent nonbonding e^-, and charges are written to identify a formal distribution of charge.

The bonds show how the atoms in a molecule are connected to each other. A Lewis structure does not show bond lengths, bond angles, the arrangement of atoms in 3D space, or the actual charges on atoms. Some molecules require more than one Lewis structure to describe them. These are called *resonance structures*. In some situations, atoms in period 3 and higher have more than $8e^-$.

Some atoms (C, N, O, and S) form double bonds, which are represented by double lines. Some atoms (C and N) can form triple bonds, represented by 3 lines.

How do you determine and draw a Lewis structure. **You have to first determine whether the compound is ionic or molecular. If it is ionic, follow Model 1. If molecular, follow Model 2.**

Critical thinking Question

1. What is the difference between a covalent bond and an ionic bond?

Types of Atoms	Type of Bond	Characteristic of Bond
Metal & Nonmetal	ionic	Complete transfer of electron from METAL to nonmetal (CHARGES)
Nonmetal & nonmetal	Molecular covalent	Sharing of electrons through a bond
Metal & Metal	metallic	Flowing of electrons throughout the metal

Model 1 Ionic Compound – complete transfer of electrons (see video)

Model 2 Lewis Structures for Molecule Compounds – sharing of electrons

Determine the central atom:

Usually, the first element written is the central atom.
Hydrogen is NEVER THE CENTRAL ATOM.
The more electronegative element is NOT the central atom.
For oxyacids (acid from an oxyanion, contains an element and O), O surrounds the central atom.
For oxyacids, the hydrogen is ALWAYS connected to the oxygen.

Step	***Example***
Tally the total number of valence electrons (VE). If it's an anion, add electrons. If its a cation, subtract electrons. The position on the PT determines the # of VE.	***nitrite***
Determine the central atom, place the other atoms around it and connect the appropriate atoms with a single bond. A single bond contains 2 e^-.	
Place lone pairs around the peripheral atoms until they achieve an octet. Hydrogens only have 2 e^-. Boron can have 6 e^-.	
Leftover electrons go on as lone pairs to the central atom.	
If the central atom is still electron deficient, convert an outer atom lone pair into a bonding pair.	

Determine the formal charge (FC) on the atoms. Evaluate whether the FC on the atoms are reasonable. FC=# of VE e^- on atom - # of lone electrons – 0.5(# of shared e^-) The structure is reasonable if the charges are zero or small, and the neg. charges reside on the most electronegative atoms.	

Things to consider and check:

- Have you used exactly the number of electrons when you added up the VE in the first step? No more, no less.
- All outer atoms have 8 electrons (H ONLY HAS 2 ELECTRONS). Occasionally, a central atom in the 3rd or higher period (extra d orbitals) may violate the octet rule (octet expansion). Except for boron, central atoms in the 2nd period and outer atoms will NEVER violate the octet rule.
- If the structure is an ion, have you enclosed the entire structure in brackets with the correct net charge outside the brackets?
- The bonding framework is often indicated by the order in which the atoms are written in a molecular formula. For example, in OCS, the carbon atom is the inner atom.
- Remember that in an oxyacid, the hydrogen atom is bonded to the oxygen, not the central atom. As an example, HNO_3, the H is not bonded to nitrogen.
- The most likely structure is the one that has the most reasonable formal charges on the atoms.
- Odd number of VE – these are radicals, use FC to determine which atom the lone e^- goes on

Three ways to count electrons (Depending on the question asked and what you are trying to determine, electrons around an atom are counted differently)

1. How many electrons are around the atom? ________________________

2. What is the formal charge? ________________________

3. What is the electron configuration? ___________________________