Atomic Structure and the Periodic Table

Evolution of Atomic Theory

- The ancient Greek scientist Democritus is often credited with developing the idea of the atom
- Democritus proposed that matter was, on the smallest scale, composed of particles described as *atomos*, meaning "indivisible"
- While atoms can in fact be broken down into smaller particles, the atoms of each element are distinct from each other, making them the fundamental unit of matter.

The Atom: A Complete Picture

- Each atom contains at its core a nucleus, a region of positive charge.
- Positively charged particles, called___________________, are contained in the nucleus.
- In addition, all atoms except hydrogen must contain one or more neutral (non-charged) particles.

  These neutral particles are called______________________

  - The protons and neutrons (the two types of particle in the nucleus) are called_________________.
.
- Which particles in an atom have negative charge?

  Where in the atom are these found?

- Note that the protons and neutrons are each almost 2,000 times more massive than an electron;

  Therefore:

- What is the approximate diameter of an atom?
- What is the approximate diameter of an nucleus?
- Therefore an atom is mostly_______________________________

Characteristics of an Atom

- The atomic number, which is symbolized by *Z*, represents the number of ___________________ in an atom.
- The number *Z* indicates which type of element that atom represents.
- These values are found in order on the periodic table.

- Example: Which element contains 5 protons?

  10 protons?

  34 protons?

- Protons are never "changed around" in a chemical reaction!
  - As we shall see, this means that an atom which begins a reaction with six protons (a carbon atom) still has six protons at the end of the reaction.
  - Using chemicals, is it possible to change lead into gold?

- Atoms that have no charge must have the same amount of positive charge as negative charge.
- Therefore, the number of electrons in an atom is equal to the number of protons in an atom (*Z*) for any *neutral* atom.
- Atoms which contain more or less electrons than protons therefore must have a charge.

  Charged atoms are called_________________.

  .

Ions: A Lesson in Thinking Backwards

- Suppose an ion has exactly one more electron than it does protons.
  - Will the ion be positively or negatively charged?

- What if an atom lost two electrons?
  - Will the ion be positively or negatively charged?

- For each electron an ion has more than it does protons, we indicate it with a – as a superscript.

- For each electron an ion has less than it does protons, we indicate it with a + as a superscript.

Examples of Ions

- A neutral bromine atom has_______protons and________electrons.

- Suppose we add exactly one electron; now we have________protons and________electrons.

- We symbolize this as ___________

- A neutral aluminum atom has_______protons and________electrons.
- Suppose the atom loses three electrons.

- We symbolize this ion as __________.

- Note that losing electrons is indicated with +, and gaining electrons is indicated with -.

Neutrons and Isotopes

- The number of neutrons in an atom *cannot* be easily predicted or found on the periodic table.
- Atoms of the same element (i.e. have the same number of protons) can have different numbers of neutrons. They are called isotopes of one another.
- Examples:
  - A carbon atom must contain 6 protons, but it may contain 6, 7, or 8 neutrons.
  - Almost all hydrogen atoms contain *zero* neutrons. The isotopes of hydrogen which contain one and two neutrons are called *deuterium* and *tritium*, respectively; they are symbolized as D and T, but do not appear on the periodic table.

More on Isotopes

- The sum of the number of protons and neutrons in an atom is called the mass number of that atom, and is symbolized by *A*.
- Isotopes are named by stating the name of the element, followed by *A*.
  - Carbon (*Z*=6) with 6 neutrons is "carbon-12"
  - Carbon with 7 neutrons is "carbon-13"
  - Carbon with 8 neutrons is "carbon-14"

- Isotopes can be written in shorthand notation in two different ways,

Examples

- For each of the following, indicate how many protons, electrons, and neutrons each atom or ion possesses.
  - ^{35}Cl
  - $^{81}Br^{-}$
  - $^{27}Al^{3+}$

The Mass of an Atom

- Recall that virtually all of the mass of an atom comes from its nucleus.
- Knowing the mass of protons and neutrons allows us to calculate the mass of one atom of a particular isotope.
- Since most elements have more than one isotope, and these isotopes are mixed in nature, it is not possible to provide an *exact* mass for the element.
- Instead, we consider a *weighted average* mass, based on the weight of each individual isotope and its abundance in nature.
- The periodic table provides the atomic weight of each element, which corresponds to this weighted average mass.
- These values are in *atomic mass units (amu)*, a very small unit of mass.
  - 1 amu = 1.66×10^{-24} g
- For example, the mass of a single helium atom is ________________________.

The Periodic Table

- Elements on the periodic table can be classified in two ways:
  - Elements in the same period are on the same row.
    - N, O, and F are in the same period.
  - Elements in the same group or family are in the same column.
    - F, Cl, and Br are in the same group.
  - Note that elements in the same group have the same number of valence electrons; therefore, they will often react in chemical reactions in a similar fashion.

Groups

- Some groups have their own name, which you must know:
  - Group 1A elements (except for hydrogen) are called the ________________________
  - Group 2A elements are called the ___________________________
  - Group 7A elements are called the ____________________________
  - Group 8A elements are called the ____________________________

Other Classifications

- The elements in Groups IA to VIIIA are collectively called the "Main group elements" or the "Representative Elements"
- The elements in the middle of the periodic table (the groups which end in B) are called "Transition metals."

Metals

- The metals include many elements found on the left side of the periodic table.
- Many metals are known for having the following properties
  - They are malleable (meaning they are soft and easily shaped)
  - They are ductile (they can be twisted and drawn into a wire)
  - They can conduct both electricity and heat.
  - They tend to be lustrous (shiny).
  - All metals are solids at room temperatures except for mercury, which is a liquid.

Nonmetals

- Nonmetals are generally found on the right side of the periodic table (except hydrogen, which is placed on the left).
- Their properties are generally the opposite of the metals.
  - Those which are solids tend to be brittle.
  - Most are poor conductors of electricity at room temperature (insulators) and do not conduct heat well.
  - Some are gases at room temperature; others are solids. Bromine is the only other element which is a liquid.

Metalloids

- Metalloids are found between the metals and the nonmetals on the periodic table.
  - They include Si, Ge, As, Sb, Te, and boron (B).
- The properties of the metalloids are often a cross between those of the metals and the nonmetals.
- All the metalloids listed are solids at room temperature.
- (I do not include Po and At with the metalloids, as they are rather unstable.)

Periodic Trends

- By comparing the position of one element to another on the periodic table, it is often possible to make comparisons between the properties of those elements; these are a result of <u>periodic trends.</u>
- We will consider one of these trends now and a second (electronegativity) later.
  - Atomic radius

Atomic Radius

- The radius of an atom tends to increase as you move down a group in the periodic table.
  - Explanation: Elements in lower groups of the periodic table have electrons in outer energy levels. These electrons are not held as tightly by the nucleus, so they can travel further away from it.
- The radius of an atom tends to decrease from left to right across a period.
  - Explanation: Electrons fill the same energy level while the positive charge in the nucleus is increasing because the number of protons is likewise increasing. This pulls the electrons in closer, making the radius smaller.

The Electromagnetic Spectrum (EM)

- EM radiation can be classified into different regions based on its wavelength or its frequency.

  - The longest wavelenths (smallest frequencies) correspond to radio waves.
  - The shortest wavelengths (greatest frequencies) correspond to gamma rays.

Electrons & Energy Levels

- Evidence suggested to scientists studying atomic structure that electrons "orbit" around the nucleus in specific <u>energy levels</u> (also called <u>shells</u>).
  - The farther the energy level is from the nucleus, the higher the energy of its electrons on average
- Normally, electrons "fill up" the lower energy levels first, and as new electrons are added they go into higher and higher energy levels.
  - An atom is said to be in the <u>ground state</u> when this is true.
- An energy source, such as light or heat, can give electrons the energy necessary to "jump" from its energy level to a higher one.
  - In this situation, the atom is said to be in an <u>excited state</u>.

Emission of Light

- An atom in the excited state is unstable, and must release the energy it gained in the first place to return to the ground state.
- One way in which this is accomplished is for the atom to give off this energy as light.
  - The released light is called a <u>quanta</u> (energy packet) or a <u>photon</u>.
- Excited electrons return to lower energy levels.

- Depending on the atom itself and which energy levels are involved, EM radiation of different wavelengths are given off.

Orbitals & Subshells

- In truth, electrons do not simply rotate about the nucleus in simple patterns.
- Instead, electrons are mostly contained in orbitals, which are shapes which describe the regions where an electron is likely to be found. Four orbital types are commonly encountered in modern chemistry.
  - These are designated *s*, *p*, *d*, and *f*, and are listed here from lowest energy orbital type to highest.
- Each individual orbital can hold, at most, two electrons in its ground state.

Orbitals and Subshells

- Each shell has a predictable number of each type of orbital
  - For example, except for the first shell, all shells possess exactly three *p* orbitals
- We classify each "set" of orbitals as a subshell
  - The three *p* orbitals in the second shell comprise the 2*p* subshell
  - The three *p* orbitals in the third shell comprise the 3*p* subshell
  - etc.

Do not confuse shells, subshells, and orbitals!

The *s* Orbital

- The *s* orbital is the least energetic of the orbitals, and, as a sphere, is also the most simple-looking.
- All energy levels have exactly one *s* orbital.
  - Since any orbital can only hold two electrons, each *s* subshell (1*s*, 2*s*, etc.) holds no more than two electrons

The *p* Orbital

- Starting with the 2^{nd} energy level, each level possesses three "sets" of *p* orbitals.
- Recall that each orbital can hold up to two electrons, so each *p* subshell will hold up to six electrons.
- The *p* orbitals are more complex in shape than the *s* orbitals, and electrons residing in them typically have greater energy than those in an *s* orbital.

***d* and *f* Orbitals**

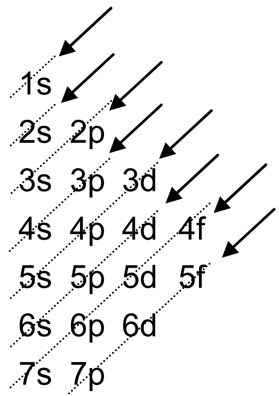
- Energy levels three and higher have a total of 5 sets of *d* orbitals.
  - A *d* subshell can therefore hold up to_____electrons.
- Energy levels four and higher have a total of 7 sets of *f* orbitals.
  - An *f* subshell can therefore hold up to_____electrons.

Electron Configurations

- Each element has its own unique electron configuration which tells us which subshells possess electrons in them and how many.
  - For example, the electron configuration of boron is $1s^2 2s^2 2p^1$, which means that there are
    - 2 electrons in an *s* subshell on the first energy level (1*s* , for short)
    - 2 electrons in the 2*s* subshell
    - 1 electron in the 2*p* subshell

Electron Configurations

- Atoms always fill their lowest energy subshells first, and successively higher ones as more electrons are added.



Filling Order of Subshells

- The order in which the electrons fill can be found on the periodic table.
  - Notice that the periodic table is broken up into 4 distinct "blocks," each of which indicates where the highest energy electrons are.

Examples

Determine the electron configuration for each of the following atoms/ions:

- Be
- N
- Na
- V
- I

Shorthand Notation for Electron Configurations

- You can abbreviate the electron configuration with a special notation
  - Consider V for example.
  - The last noble gas before V was Ar (element 18)
  - So, we can write [Ar]$4s^2 3d^3$ as vanadium's electron configuration.
  - What is the electron configuration of antimony in shorthand notation?

Valence Electrons

- The electrons occupying the outermost energy level of an atom are called the valence electrons; all other electrons are called the core electrons.
- The valence electrons, as we will see, are responsible for chemical bonding.
  - Knowing the number of valence electrons an atom has is the single most important information you can have in predicting how an atom will react chemically.
- Note that the valence electrons are not *always* the electrons with the highest energy, as we will see with the transition metals

- For elements in group IA through VIIIA, the number of valence electrons an atom has is simply its group number.
  - So phosphorus, in group VA, has five valence electrons.
- Helium is the only significant exception; it has 2 valence electrons, even though it is in group VIIIA

Standard States of the Elements

This periodic table shows you the state (gas, liquid, or solid) of many important elements at room temperature. Notice that all metals and metalloids<u>except mercury</u> are solids at room temperature. The only two elements which are found as liquids are bromine and mercury. All the elements in the last group (8A) are gases. The rest are fairly easy to memorize. On this periodic table I have indicated the diatomic elements with their formulas; you will not see this on a "normal" periodic table.

Key: solid | liquid | gas

Group 1A	2A	3B	4B	5B	6B	7B	8B			1B	2B	3A	4A	5A	6A	7A	8A
H_2 Hydrogen																	**He** Helium
Li Lithium	**Be** Beryllium											**B** Boron	**C** Carbon	**N_2** Nitrogen	**O_2** Oxygen	**F_2** Fluorine	**Ne** Neon
Na Sodium	**Mg** Magnesium											**Al** Aluminum	**Si** Silicon	**P** Phosphorus	**S** Sulfur	**Cl_2** Chlorine	**Ar** Argon
K Potassium	**Ca** Calcium	**Sc** Scandium	**Ti** Titanium	**V** Vanadium	**Cr** Chromium	**Mn** Manganese	**Fe** Iron	**Co** Cobalt	**Ni** Nickel	**Cu** Copper	**Zn** Zinc	**Ga** Gallium	**Ge** Germanium	**As** Arsenic	**Se** Selenium	**Br_2** Bromine	**Kr** Krypton
Rb Rubidium	**Sr** Strontium								**Pd** Palladium	**Ag** Silver	**Cd** Cadmium		**Sn** Tin	**Sb** Antimony		**I_2** Iodine	**Xe** Xenon
Cs Cesium	**Ba** Barium				**W** Tungsten				**Pt** Platinum	**Au** Gold	**Hg** Mercury		**Pb** Lead	**Bi** Bismuth			

		U Uranium											

Orbital Blocks