Chemistry -

- Elements
- Electron Configurations
- The Periodic Table

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History of Chemistry

Before 16th Century

Alchemy – Attempts (scientific or otherwise) to change cheap metals into gold – no real science

17th Century

Robert Boyle: First "chemist" to perform quantitative experiments

18th Century

- **Many elements discovered such as oxygen**
- **Law of Conservation of Matter discovered – "Matter is neither created or destroyed in chemical changes" Lavoisier**

Dalton's Atomic Theory

1800 - First Comprehensive Atomic Theory

1. **Matter is composed of atoms that are indivisible**
2. **All atoms of the same element are identical**
3. **Atoms combine in small whole number ratios to form compounds**
4. **Atoms are not created or destroyed in chemical reactions - Conservation of Matter - Lavoisier**

Now let's define elements and compounds

Element – matter composed of 1 type of atom

Compound – matter where atoms of different elements are bonded to each other

History of Periodic Table

Mendeleev - 1850

- **Arranged elements by atomic mass in periodic table**

Moseley - 1915

- **Discovered how to find atomic number of elements**
- **Arranged elements by atomic number - this is modern periodic law**

The arrangement of elements is by reactivity and structure. Reactivity was known before structure was known.

The Structure of the Table

Periods - horizontal rows, numbered from top to bottom

- **Be able to identify lanthanide and actinide series of elements. Collectively they both belong to the "rare earths".**

Groups - vertical columns, numbered in several ways: using A and B designations or by numbers

Names of Groups to know -

- **Alkali Metals – IA or column 1**
- **Alkaline Earth Metals – IIA or column 2**
- **Transition Metals - IIIB – IIB or columns 3-12**
- **Chalcogens - VI A or column 16**
- **Halogens – VIIA or column 17**
- **Noble Gases – VIIIA or column 18**

Another way to divide the table

- **Metals - have a tendency to lose electrons in bonding, good electrical and thermal conductivity; elements on the left side of the table**
- **Nonmetals - have a tendency to gain electrons in bonding, usually have poor electrical and thermal conductivity; elements on the upper right of the table**
- **Metalloids - have properties between metals and nonmetals, reside along the “stairstep”**

The Elements

A. Atoms

Electron	**9.11×10^{-28} g**	**-1 charge**
Proton	**1.673×10^{-24} g**	**+1 charge**
Neutron	**1.675×10^{-24}g**	**neutral**

Radius of atom	**10^{-10} m**
Radius of nucleus	**10^{-14} m**
Radius of proton & neutron	**10^{-15} m**
Radius of electron	**10^{-18} m**

The concept of atomic weight or average atomic mass

Atomic number - number of protons (identifies the element)

Mass number - number of protons plus neutrons

$$^{\text{Mass number}}_{\text{Atomic Number}}\mathbf{X}$$

Atomic mass unit - 1/12 the mass of a carbon 12 atom (1.66 x 10^{-24} g)

Each atom has an atomic mass based on the amu (atomic mass unit)

Isotopes - different forms of an element, same atomic number but different mass numbers

Atomic weight - the weighted average (based on relative abundance) of the atomic masses of the isotopes of an element

$$\text{Atomic weight} = \frac{\text{mass of a given number of atoms}}{\text{given number of atoms}}$$

Electron Configurations

The key to chemistry is understanding electrons and the new ideas from the Schrodinger "quantum mechanical model".

The "quantum mechanical model" tells us that electrons do not just reside in simple circular orbits as in the Bohr model.

The structure is more complex and each electron is different from another electron by its "quantum numbers" – 4 numbers needed to solve the wave equation for the electron.

letter	**Name**	**Property**	**Values**
n	**principal quantum number**	**energy level**	**1,2,3....**
l	**angular momentum**	**energy sublevel**	**0 to (n-1)**
m_l	**momentum in z direction**	**# of orbitals in sublevel**	**-l to +l**
m_s	**spin**	**# of electrons in orbital**	**$\pm$1/2**

These numbers tell us that there are energy sublevels in each energy level and within each energy sublevel are orbitals (electron houses). Only 2 electrons can be in each orbital and those electrons must have the opposite spin. The situation is much more complex than the Bohr model!

Spectroscopic notation

When writing electron addresses chemists aften use a shorthand notation called spectroscopic notation for electron configurations. In this notation the l quantum number values are renamed as follows

l = 0	s sublevel	1 orbital	2 electrons
l = 1	p sublevel	3 orbitals	6 electrons
l = 2	d sublevel	5 orbitals	10 electrons
l = 3	f sublevel	7 orbitals	14 electrons

This allows for a condensed form of notation. For example the 6 electrons in energy level 2 and the p sublevel would be referred to as $2p^6$

The filling of electron orbitals

To predict how many electrons will be in each energy level and sublevel we need to know the energies of electron orbitals. Due to the increasing closeness of the energy levels **and the** sublevel splitting of the energy level **the energy sublevels from one level start to overlap the sublevels of the next energy level at n = 3.**

We will use as a general rule the idea that electrons will fill the lowest energy orbitals possible**. As a result of the overlap described above the electrons fill the orbitals of energy level 1 and 2 completely but do not fill energy level 3 completely until part of energy level 4 is filled.**

The Periodic Table is an invaluable aid in determining the order of sublevel filling.

Remember:

- blocks of two will correspond to the filling of the s sublevel (alkali and alkaline earth metals)
- blocks of 6 the p sublevel (this includes the halogens and noble gases)
- blocks of 10 the d sublevel (the transition metals)
- blocks of 14 the f sublevel. (the lanthanide and actinide series)

There are exceptions, most notably:

1. **Major**
   - **Cr and Mo are s^1d^5 and not s^2d^4.**
   - **Cu, Ag and Au are s^1d^{10} not s^2d^9. This can be explained by the extra stability of a full sublevel.**
2. **Minor**
   - **Nb, Ru, Rh, Pd, Pt**

- **Many of the Lanthanide and Actinide series**

General Trends in the Periodic Table

As we go down a family

1. The outermost electrons are farther from the nucleus
2. The inner electrons shield the outermost ones from the nucleus so the effective nuclear charge decreases.
3. The outermost electrons are held less tightly and the atom gets larger.
4. Reactivity as a metal increases.

As we go across a period

1. The outermost electrons are about the same distance from the nucleus
2. There is very little additional shielding so the effective nuclear charge increases.
3. The outermost electrons are held more tightly and the atom gets smaller.
4. Reactivity as a nonmetal increases.

Summary of trends

1. Size increases as we go down a family and decreases across a period
2. Ionization energy increases as we go across a period and decreases as we go down a family.
3. Metallic behavior increases diagonally toward francium at the bottom left. Nonmetallic behavior increases diagonally toward fluorine at the top right. Metal ions are smaller than metal atoms. Nonmetal ions are larger than nonmetal atoms.
4. Electronegativity (attraction for additional electrons in a bonding situation) increases as we go across and decreases as we go down.

Prediction of Oxidation numbers

The driving force for elements to lose or gain electrons can be calculated from quantum theory in a very complex way. The results can be summarized very simply:

Elements would like to have 8 electrons in their outermost energy level (highest n value). This means that elements try to get a s^2p^6 configuration in their highest n value energy level. This is the octet rule.

The oxidation number of an element is its tendency to lose or gain a certain number of electrons in a bonding situation. It is a book-keeping system which is useful in writing formulas and naming compounds. In ionic bonding it coresponds to the charge on the ion. In a covalent bond it does not represent the real charge but is still useful.

Some general rules for predicting oxidation numbers are:

1. Alkali metals +1
2. Alkaline earth metals +2
3. Transition elements - lose s and then some or all of the d electrons, difficult to predict
4. Boron family +3 but Tl is +1, +3
5. Carbon family usually +4 but Sn and Pb are +2,+4
6. N family is -3 but all can also exhibit +3, +5 in nonmetal compounds
7. O family is -2 to Po; these elements can also exhibit +4 and +6 in nonmetal compounds
8. Halogens are -1, but in compounds with nonmetals can exhibit +5 and + 7
9. Lanthanide and Actinide series - most common is +3 but others are possible

<u>You will not be required to memorize a list of oxidation numbers!</u>