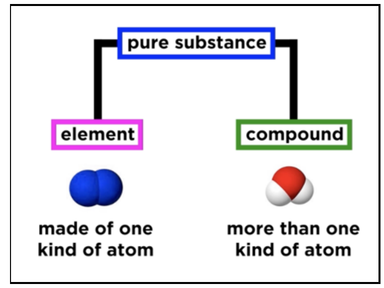
AP Chemistry Ultimate Review Packet
Unit 1: Atomic Structure and Properties

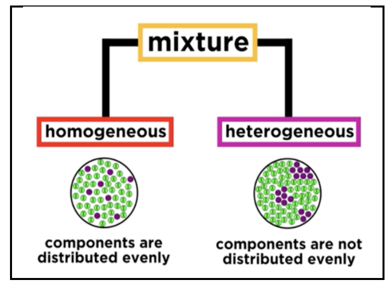
Atoms and Molecules, Pure Substances and Mixtures



Atoms are the smallest unit of an element (substances on the [periodic table])

Molecules contain two or more atoms that are chemically bound:
- molecules can be **elements** if they contain one kind of atom
- molecules can be **compounds** if they contain more than one kind of atom

Pure substances include elements and compounds



Mixtures contain two or more pure substances
- mixtures can be **homogeneous** if substances are dispersed *evenly*
- mixtures can be **heterogeneous** if substances are dispersed *unevenly*

The Structure of the Atom

Atoms contain **protons (+)**, **neutrons**, and **electrons (-)**

Atomic number = number of protons (determines the element the atom belongs to)

Mass number = number of protons + number of neutrons

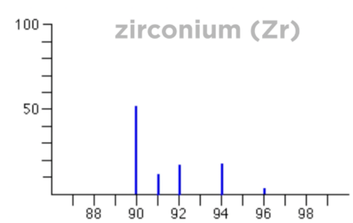
Atoms of a given element can have different mass numbers if they have differing numbers of **neutrons** (these are called **isotopes**)

Atomic mass of an element is the average mass of an atom of that element, or the average of the mass numbers of its naturally occurring isotopes with respect to their relative abundance

Electrons are of negligible mass compared to protons and neutrons, which have roughly equivalent mass

Atoms can have formal electric charge, in which case they are called **ions**
- atoms with equal numbers of electrons and protons are of **neutral charge**
- atoms with more electrons than protons have a **negative charge** (anion)
- atoms with fewer electrons than protons have a **positive charge** (cation)

Isotopic masses and abundances can be determined by **mass spectrometry**
- sample is vaporized and ionized and sent through a tube with a magnetic field
- particles separate based on mass to charge ratio (m/z)



Mass numbers on the horizontal axis, **relative abundance** (%) on the vertical axis

$^{24}_{12}\mathrm{Mg}^{2+}$

We can represent atoms with nuclide symbols:
- left subscript is **atomic number** (# of protons)
- left superscript is **mass number** (# of protons + # of neutrons)
- right superscript is **electric charge** (# of protons - # of electrons)

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Avogadro's Number and the Mole

When we want to discuss macroscopic quantities of matter we need to use **moles**

- 1 mole = Avogadro's number = 6.022×10^{23}
- converts between atomic mass units and grams
  - 1 carbon atom = 12 amu, so 1 mole of carbon atoms = 12 grams

Elements have a **molar mass** equal to their atomic mass, just in g/mol instead of amu
Compounds have a **molar mass** equal to the sum of the atomic masses of all the atoms present within the compound

- example: glycine ($C_2H_5O_2N$)
  C: (2)(12) = 24
  H: (5)(1) = 5
  O: (2)(16) = 32
  N: (1)(14) = 14

24 + 5 + 32 + 14 = 75 so glycine has a molar mass of 75 g/mol

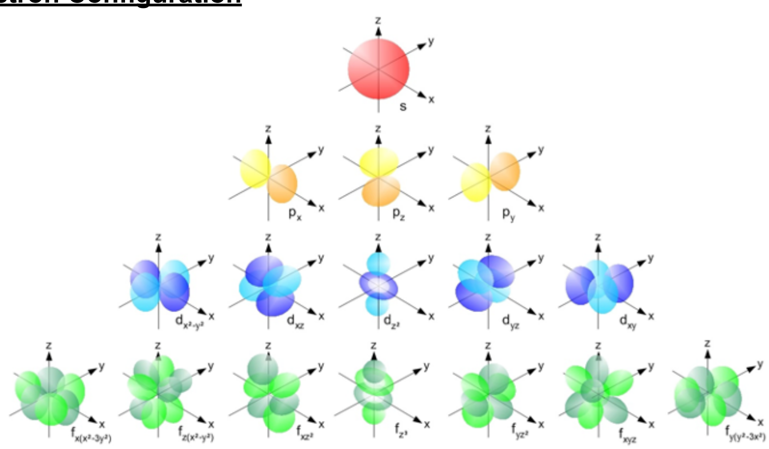
To convert moles into mass: multiply by the molar mass
To convert mass into moles: divide by the molar mass

Percent composition: percent of molecular mass represented by each element
NH_3: N = 14/17 = 0.82 = 82%, H = 3/17 = 0.18 = 18%
Empirical formula: lowest whole number ratio of elements in a compound. Find via elemental analysis:
1) combust unknown sample and collect masses of C and H
2) convert masses into moles
3) divide both molar values by the smaller to get a ratio
4) express empirical formula with that ratio (e.g. 1:2 = CH_2)

Molecular formula: actual number of atoms of each element in a compound. Take empirical formula and find how many multiples are required to reach the molecular mass determined by mass spectrometry, multiply through subscripts

(CH_2 = 14, MM = 42, 42/14 = 3, so MF = C_3H_6)

Atomic Orbitals and Electron Configuration



Electrons inhabit **atomic orbitals** which can be described by **quantum numbers**
1) n = principal quantum number (energy level)
2) ℓ = angular momentum quantum number (type of orbital)
   $\ell = 0 \rightarrow$ s orbital (1 per energy level)
   $\ell = 1 \rightarrow$ p orbitals (3 per energy level)
   $\ell = 2 \rightarrow$ d orbitals (5 per energy level)
3) m_ℓ = magnetic quantum number (individual orbital within subshell)
4) m_s = spin quantum number (+1/2 or -1/2)

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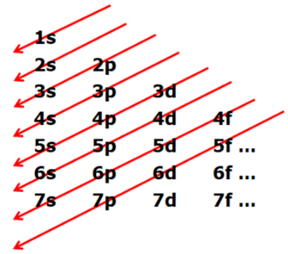
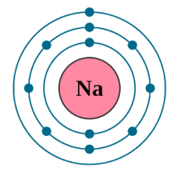
Electrons fill up these orbitals (max. two per orbital) according to:
1. **Aufbau principle** – tells us the order in which orbitals are filled
2. **Hund's rule** – says that subshells are half-filled before doubling any orbital
3. **Pauli exclusion principle** – no two electrons within an atom can have precisely the same of quantum numbers



Electron configurations tell us which orbitals are inhabited by the electrons in an atom:

e.g. $1s^2 2s^2 2p^6 3s^1$ → two electrons in the 1s orbital (full)
two electrons in the 2s orbital (full)
six electrons in the 2p orbitals (full)
one electron in the 3s orbital (half-full)

- the diagram shows the ground state electron configuration of **sodium** (Na)



The Periodic Table and its Trends

Rows of the periodic table are called **periods.** Columns of the periodic table are called **groups.**
- elements in the same group have similar electron configurations
- they have the same number of **valence electrons** (outermost shell)

The periodic table displays important **periodic trends:**

1) covalent radius: increases down and to the left
- essentially describes the size of an atom
- going down we add shells (gets bigger with each shell)
- going right we add protons (more attraction causes contraction)
- cations are always smaller than neutral atom, anions always larger

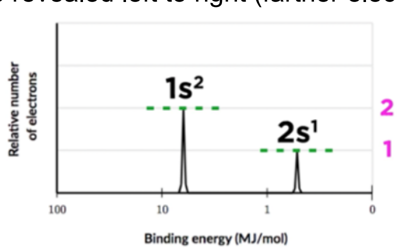
2) ionization energy: increases up and to the right
- energy required to remove the outermost electron from ground state
- opposite of radius because closer to nucleus = harder to ionize
- each successive ionization energy gets larger

3) electron affinity: increases up and to the right
- refers to the energy associated with neutral atom gaining an electron
- harder to ionize means more likely to gain an electron
- could involve absorbing or releasing energy
- noble gases are omitted from the trend as they have a full valence shell

4) electronegativity: increases up and to the right
- a measure of how tightly an atom holds electron density
- smaller atoms have electrons closer to the nucleus and attract better
- some noble gases are omitted from the trend as they do not form bonds

Photoelectron spectroscopy reveals the electron configuration of an atom
- horizontal axis displays binding energy (related to ionization energy)
- vertical axis displays number of electrons
- Aufbau principle revealed left to right (farther electrons easy to ionize)

lithium
$1s^2 2s^1$



oxygen
$1s^2 2s^2 2p^4$