Chemistry 1314 – Lecture Outline and Notes
Text: Chemistry – The Central Science, 10^{th} edition, Brown, LeMay and Bursten

Chapter 2. Atoms, Molecules, and Ions

Atomic theory (postulated by John Dalton, 1803-1807):
1. Elements composed of atoms
2. All atoms of an element are identical; and each element composed of unique atoms
3. Atoms are neither created nor destroyed in chemical reactions (Law of Conservation of Matter)
4. Compounds formed by combination of atoms of elements – a given compound always has the same relative number and kind of atoms (Law of Constant Composition)

The Law of Multiple Proportions is derived from the atomic theory. Two elements can combine into different compounds, but their ratio is always a whole number.

Atomic Structure
Atoms composed of 3 fundamental particles: electrons, protons, and neutrons. Protons and neutrons are together in the nucleus with the electrons located outside the nucleus.

ELECTRON:
- Discovered by JJ Thomson in 1897
- Cathode Rays are a stream of negatively charged particles (electrons)
- Obtained the charge-to-mass ratio of the cathode rays (1.76×10^{8} C/g), and this value is independent of the material used in the cathode ray tube
- In 1909, Robert Millikan determined the charge of the electron (-1.6022×10^{-19} C) (C is Coulombs).
- Using the charge-to-mass ratio and the charge of the electron, the mass of the electron can be determined
    - Electron mass = 1.60×10^{-19} C/1.76×10^{8} C/g = 9.10×10^{-28} g

PROTON:
- Thomson had detected positively charged particles in the cathode ray tube, but not able to characterize
- In 1919, Ernest Rutherford determined that the proton has a charge equal, but opposite, to that of the electron. The mass was measured as 1.673×10^{-24} g.

NEUTRON:
- Characterized by James Chadwick in 1932
- Has almost the same mass as the proton, but no charge

Particle	Charge	Mass	Relative charge
Electron	-1.6022×10^{-19} C	9.10939×10^{-28} g (5.486×10^{-4} amu)	1-
Proton	$+1.6022 \times 10^{-19}$ C	1.673×10^{-24} g (1.0073 amu)	1+
Neutron	0	1.675×10^{-24} g (1.0087 amu)	0

Current Atomic Model:

Rutherford performed an experiment where he shot α-particles (essentially the nuclear core of helium) at thin gold foil. The scatter of the particles led Rutherford to develop a new model of the atom, the one that we currently hold. The nucleus contains most of the mass and is composed of the protons and neutrons. The electrons are outside the nucleus.

- Atoms are on the size of 100-500 pm (or 0.1-0.5 nm). A useful unit is the angstrom (Å), defined as 10^{-10} m.

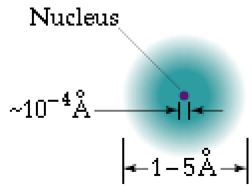


Figure 2.11. Schematic cross-sectional view through the center of an atom. Virtually all of the mass of an atom resides in the nucleus, which contains protons and neutrons. The rest of the atom is the space in which the light, negatively charged electrons reside.

<u>Atomic Number</u>

How many electrons, protons, and neutrons does each atom contain? It depends on the element.

- Each element has only one kind of atom
- The periodic table can tell us the number of electrons, protons, and neutrons.
- NOTE: Moseley discovered a correlation between X-rays released by an atom and its number of protons. He proposed that the elements can be ordered by using an "atomic number" – a whole number representing the number of protons in an atom.
- The number above the element symbol is the atomic number and equals the number of protons for that element
- The number of electrons equals the number of protons for a neutral atom

Examples:

1. Carbon – symbol ___, atomic number =___, thus ___ protons and ___ electrons
2. Iron – symbol ___, atom number = ___, thus ___ protons and ___ electrons
3. Mercury – symbol ___, atomic number = ___, thus ___ protons and ___ electrons
4. Uranium – symbol ___, atomic number = ___, thus ___ protons and ___ electrons

What about the number of neutrons – we first must understand the atomic mass.

<u>Atomic Mass</u>

- The number below the element symbol is the atomic mass. For hydrogen, it is ________
- We think atomic mass is the mass of the atom, but a hydrogen atom has a mass of 1.6735 x 10^{-24} g. Why the difference?
- We note that the atomic mass does not contain a unit. It is a "relative" mass.
- The atomic mass is relative to a carbon atom composed of 6 protons, 6 electrons, and 6 neutrons. This particular carbon has a mass of 1.99268 x 10^{-23} g. Therefore, 1/12 of the mass of this carbon is defined as an atomic mass unit (amu or SI u). It is also called a Dalton (Da).

- $1/12 * 1.99268 \times 10^{-23}$g = 1.66057×10^{-24}g = 1 amu
- Mass of a hydrogen atom:
  - 1.6735×10^{-24}g (1 amu/1.66057×10^{-24}g) = ______________

Isotopes:

But the atomic mass of hydrogen is 1.00797 amu. Why does it differ? The reason for the differences is that a sample of an element is not composed of exactly the same type of atom. A pure sample of an element will have atoms composed of the same number of protons, but might contain different numbers of neutrons. Therefore, a pure sample can have atoms of different mass. The atoms that differ in the number of neutrons are called isotopes.

- Hydrogen has two different isotopes
  - One has 1 proton and 0 neutrons
  - The other has 1 proton and 1 neutron – it is also called deuterium
- Carbon has two stable isotopes ^{12}C and ^{13}C, and many unstable ones. The most common being ^{14}C. An atom of a specific isotope is called a nuclide (^{14}C is referred to as ^{14}C nuclide).

Relative, weighted average atomic mass

The atomic mass found in the periodic table accounts for the stable isotopes and their masses to obtain a weighted average atomic mass. “Weighted Average” means we must account for the differences in the abundances of the isotopes.

Symbol	Atomic Mass (amu)	Abundance
^{1}H	1.0078250	99.985%
^{2}H	2.0141017	0.015%

So, a pure sample of hydrogen contains mostly ^{1}H with a very small amount of ^{2}H. To calculate the weighted atomic mass we can not simple take the average. We must take the abundance into account.

Weighted average = (atomic mass * fractional abundance)$_1$ + (atomic mass * fractional abundance)$_2$ + … + (atomic mass * fractional abundance)$_n$

Therefore, for hydrogen:
Weighted average =

– a match to the periodic table!!!

The atomic mass in the periodic table is a relative, weighted average atomic mass. If you round the atomic mass to the nearest whole number, it is equal to the number of protons and the average number of neutrons per atom of the element.

Naturally occurring carbon is composed of 98.93% ^{12}C (12 amu (exactly)) and 1.07% ^{13}C (13.00335 amu).
Average atomic mass =

Note: The average atomic mass is also known as the <u>atomic weight</u>

Example of determining the particles from isotopes:

- ^{109}Ag – protons = ____, electrons = ____, neutrons = ___
- ^{24}Al – protons = ____, electrons = ____, neutrons = ____

<u>The Periodic Table</u>

Very important tool for the chemist and contains a large quantity of information.
We have discussed the element symbol, the atomic number and atomic mass (or atomic weight).
The elements are arranged by their atomic number, and grouped in columns by common physical properties. The physical properties show a repeating, or <u>periodic</u>, pattern.

Three groups for the elements:
METALS – left side **except hydrogen**

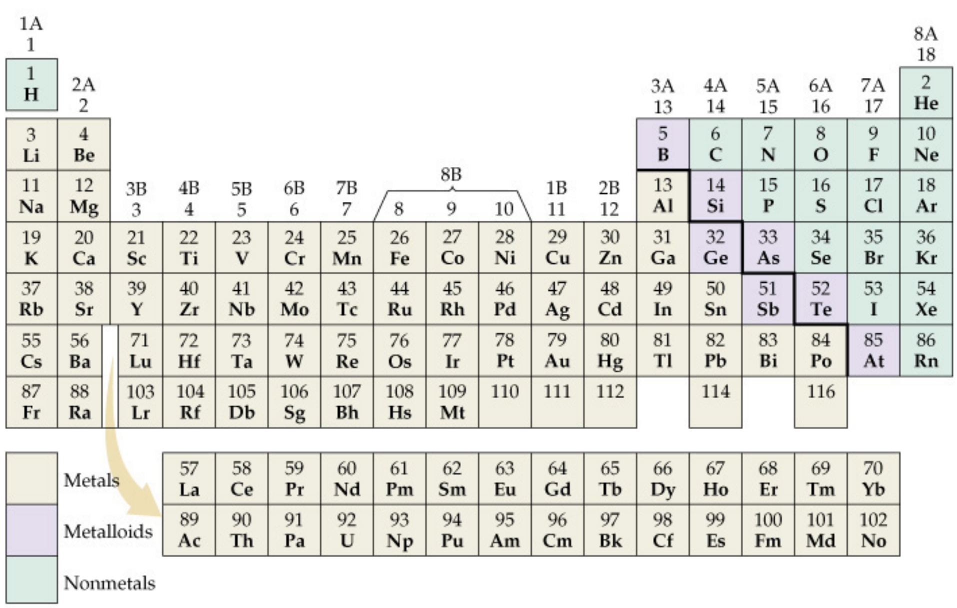
- All solids (except mercury)
- Lustrous appearance
- Malleable (able to be hammered into thin sheets) and ductile (able to be drawn into a wire)
- Good conductors of heat and electricity

NONMETALS

- Are found as gas, liquid, or solid
- Lack the remaining properties of the metals (not malleable and ductile, and are usually poor conductors (graphite (form of carbon) is a very good electrical conductor))

METALLOIDS

- Border the division of metals and nonmetals and share properties from both groups.
- B (boron), Si (silicon), Ge (germanium), As (arsenic), Sb (antimony), Te (tellurium), and At (astatine)



Molecules and Molecular Compounds

Molecule – 2 or more atoms tightly bound together.

Some elements exist as a pair, called diatomic molecules. These are hydrogen (H_2), oxygen (O_2), nitrogen (N_2), and the halogens: fluorine (F_2), chlorine (Cl_2), bromine (Br_2), and iodine (I_2).

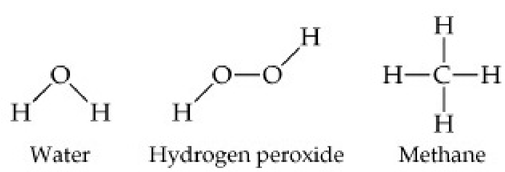
Oxygen also appears as ozone (O_3), which has very different chemical and physical properties from O_2

Molecular compounds – compounds that contain more than one type of atom. Example water represented by the **chemical formula** H_2O, methane CH_4, carbon dioxide CO_2

Molecular formula – tells the actual numbers and types of atoms in the molecule
Empirical formula – tells just the types and the ratio of the atoms in the molecule
Example: H_2O_2 is the molecular formula for hydrogen peroxide, but its empirical formula is HO.

Certain experiments only provide the types of atoms and their ratios, but not the actual molecular formula. Also, some molecules are complex and the empirical formula is the best way to describe them, for example graphite is just C for carbon.

Structural formula – shows how the atoms are connected, but doesn't really show the proper angles. Examples:



Ions and Ionic Compounds

Ion – formed when an atom gains or looses an electron

- Cation – positive charge (loss of an electron)
- Anion – negative charge (gain of an electron)
  - Examples Na^+ (cation) and Cl^- (anion)

In general, metal atoms tend to loose electrons to form cations, whereas nonmetal atoms tend to gain electrons to form anions.

Example: Give the chemical symbol, including mass numbers, for the following ions:
The ion with 26 protons, 30 neutrons, 23 electrons ___________________

Simple ions – single atoms, such as Na^+ and Cl^-
Polyatomic ions – atoms joined as in a molecule but with a charge
Examples, NO_3^- (nitrate ion) and SO_4^{2-} (sulfate ion)

Predicting Ionic Charge

Some elements we can predict their preferred charge – those that try to "mimic" the noble gases. Examples are the alkali metals group 1A elements (loose an electron; Na^{+}), the alkaline earth metals group 2A elements (loose 2 electrons; Mg^{2+}), and the halogens group 7A (gain an electron, Cl^{-})

Ionic compounds – compounds that contain both positively and negatively charged ions.
Example: table salt, NaCl.
Ionic compounds are represented by an empirical formula

A quick way to write the empirical formula for an ionic compound is by examining the charge on the ions being used. For example, the ionic compound formed by Mg (which forms Mg^{2+} ions) and N (which forms N^{3-} ions) is Mg_3N_2 (magnesium nitride).

Mg^{2+} N^{3-} -> Mg_3N_2 (cross the charges).

<u>Naming Inorganic Compounds</u>

Called **chemical nomenclature**

<u>Rules of nomenclature:</u>

- Major division
  - *organic compounds* (contain carbon)
  - *inorganic compounds* (all others)

<u>Names and Formulas of Ionic Compounds</u>

Positive Ions (Cations) – the metal portion

- Cations formed from metal atoms have the same name as the metal
  - Na^{+} - sodium ion
  - Zn^{2+} - zinc ion
  - Ba^{2+} - barium ion

Ions formed from a single atom are called monoatomic ions

- If a metal can form cations of differing charge, the positive charge is given by a Roman numeral in parentheses following the name of the metal
  - Fe^{2+} - iron(II) ion Fe^{3+} - iron(III) ion
  - Cu^{+} - copper(I) ion Cu^{2+} - copper(II) ion
  - These ions display different properties.
  - Most metals with variable charge are *transition metals* (middle of periodic table)
  - Metals without variable charge are group IA, group 2A, Al^{3+} in group 3A, and two transition metals: Ag^{+} and Zn^{2+}. It is not incorrect to indicate the charge with Roman numerals for these species.
  - Older method, but still seen and should be recognized. Add *–ous* or *–ic* to the ending of the Latin name for the metals for the lower and higher charge, respectively.
    - Fe^{2+} - ferrous ion Fe^{3+} - ferric ion
    - Cu^{+} - cuprous ion Cu^{2+} - cupric ion

- Cations formed from nonmetal atoms have names that end in *–ium*.
    - NH_4^+ - ammonium ion H_3O^+ - hydronium ion
    - These examples are polyatomic (composed of many atoms)

NOTE Table 2.4 (p. 62) for names of common cations

Hg_2^{2+} is odd because it is polyatomic, but is named mercury(I) or mercurous ion. The other mercury species is Hg^{2+}, called mercury(II) or mercuric ion.

Negative Ions (Anions)

- Monotonic anions are made by replacing the ending of the element name with *–ide*
    - H^- - hydride ion O^{2-} - oxide ion N^{3-} - nitride ion
    - A few polyatomic also follow this rule
        - OH^- - hydroxide ion CN^- - cyanide ion O_2^{2-} - peroxide ion
- Polyatomic anions containing oxygen (called oxyanions) have names ending in *–ate* or *–ite*. –ate is used for the most common oxyanion element and –ite is used for an oxyanion that has the same charge but one less oxygen.
    - NO_3^- - nitrate ion NO_2^- - nitrite ion
    - SO_4^{2-} - sulfate ion SO_3^{2-} - sulfite ion
    - We add prefixes when there are more than 2 oxyanions. *per*- is added for one more O than the oxyanion with the –ate. *hypo*- is one less oxygen than the –ite oxyanion.
        - ClO_4^- - perchlorate ion
        - ClO_3^- - chlorate ion
        - ClO_2^- - chlorite ion
        - ClO^- - hypochlorite ion

If you learn the most common oxyanion, then you can determine the name for the other oxyanions in that series.

- Anions derived by adding H^+ to an oxyanion are named by adding the word *hydrogen* or *dihydrogen* as a prefix
    - CO_3^{2-} - carbonate ion HCO_3^- - hydrogen carbonate ion
    - PO_4^{3-} - phosphate ion $H_2PO_4^-$ - dihydrogen phosphate ion
    - Note: the addition of H^+ reduces the charge on the anion.
    - Older method is to use the prefix *bi*-
        - HCO_3^- becomes bicarbonate ion
        - HSO_4^- becomes bisulfate ion
    - Names and formulas of common anions in Table 2.5 (p. 64)

Ionic Compounds

Consists of the cationic name followed by the anion name:

$CaCl_2$ (composed of Ca^{2+} and Cl^-) is *calcium chloride*

$Al(NO_3)_3$ (composed of Al^{3+} and NO_3^-) is *aluminum nitrate*

$Cu(ClO_4)_2$ (composed of Cu^{2+} and ClO_4^-) is *copper(II) perchlorate* (or *cupric perchlorate).*

Examples: Name the following
1. K_2SO_4 –
2. $Ba(OH)_2$ –
3. $FeCl_3$ –

Examples: Write the formulas for the following compounds
1. magnesium sulfate –
2. silver sulfide –
3. lead(II) nitrate –

Naming and Formulas of Acids
Acids are an important class of hydrogen-containing compounds with a special way of naming. For our current purposes, an acid yields H^+ when dissolved in water, and will have H as the first element, as in HCl and H_2SO_4. There are enough H^+ to neutralize the anion. The naming is related to the name of the anion.

- Acids based on anions whose names end in *–ide*. They have the prefix *hydro-* and an *–ic* ending.
  - Cl^- (chloride) HCl (hydrochloric acid)
  - S^{2-} (sulfide) H_2S (hydrosulfuric acid)
- Acids based on anions whose names end in *–ate* have associated with it an *–ic* ending, and those ending with *–ite* have associated with it an *–ous* ending.
  - ClO_4^- (perchlorate) $HClO_4$ (perchloric acid)
  - ClO_3^- (chlorate) $HClO_3$ (chloric acid)
  - ClO_2^- (chlorite) $HClO_2$ (chlorous acid
  - ClO^- (hypochlorite) HClO (hypochlorous acid)

Names and Formulas of Binary Molecular Compounds
The naming of binary (2 element) compounds is similar to that for ionic compounds

- The name of the element farthest to the left in the periodic chart is usually written first. EXCEPTION: compounds that contain oxygen, where oxygen is written last except when combined with fluorine.
- If both elements are in the same group, the lower one is named first
- The name of the second element is given an *–ide* ending
- Greek prefixes (Table 2.6 p. 66) are used to indicate the number of atoms of each element. NOTE: *mono-* is never used for the first element, also if the prefix ends in a or o and the name of the second element begins with a vowel, the a or o is often dropped.
  - Cl_2O dichlorine monoxide
  - NF_3 nitrogen trifluoride
  - N_2O_4 dinitrogen tetroxide
  - P_4S_{10} tetraphosphorus decasulfide

Some Simple Organic Compounds
Organic chemistry involves the study of compounds that contain carbon.

- Alkanes – contain only carbon and hydrogen are called **hydrocarbons**. When each carbon is bound to 4 other atoms, it is called an **alkane**.
  - methane (CH_4)
  - ethane (C_2H_6)
  - propane (C_3H_8)
  - Each ends in *–ane*.
  - Can continue to add carbons to the chain, the naming is derived from the Greek prefixes. Example – octane (C_8H_{18}).

- Some derivatives of alkanes – obtained from replacement of hydrogen atoms with *functional groups*. An alcohol, for example, is obtained from replacing an H atom with an –OH group. The name is obtained by adding *–ol* to the ending.
  - methanol
  - ethanol
  - 1-propanol (The one indicates the placement of the –OH group)
  - Carbon containing compounds can obtain varying forms and complexity. You will learn more in later chapters, and many of you will endure a year of organic chemistry.