Chemistry Chpt 2 (08/27/2018)

Elements and Compounds :

Element:

- cannot be decomposed into a simpler substance through chemical processes; distinguished by the unit of the atom

Compound:

- a substance made from the atoms of two or more elements bonded chemically in defined proportions
- Compounds can only be decomposed into their respective elements via chemical processes

Pure Substances and Mixtures :

Pure substance:

- A substance with a fixed and uniform composition and distinct properties (ex: pure water).

Mixture:

- A combination of two or more pure substances which can vary in composition and properties
    a) homogeneous example: salt water
    b) heterogeneous example: oil and water
- It is possible to separate mixtures through physical processes

Early Chemical Discoveries

Lavoisier (1743 – 1794) :

- Conservation of Matter : A chemical equation must be balanced for mass.

Dalton (1766 – 1844) :

- Atomic Theory : A chemical equation must have the same number of atoms of the same kind on both sides.

Law of Conservation of Mass

- The total mass during a physical or chemical process must remain constant

Ex. $2AgNO_{3\,(aq)} + Na_2CrO_{4\,(aq)} \rightarrow 2NaNO_{3\,(aq)} + Ag_2CrO_{4\,(s)}$

Law of Definite Composition

- All samples of a given compound have the same composition.

Law of Multiple Proportions

- Combinations of elements are in ratios of small whole numbers (to make molecules!)

  Carbon monoxide (CO):

  - 1.33 g of oxygen combines with 1.0 g of carbon or O/C = 1.33

  Carbon dioxide (CO_2):

  - the ratio of O/C is 2.667

  2.667/1.33 = 2.0

  ...which is the ratio of oxygen in CO_2/oxygen in CO

Development of Dalton's Theory

- Dalton's work helped to explain several empirical observations:
  - the law of constant composition
  - the law of multiple proportions
  - the law of conservation of mass

Dalton's Atomic Theory

1. All matter is made of small, indestructible particles called "atoms"
2. All atoms of a given element are identical and atoms differ for different elements
3. Atoms of one element cannot be transformed into atoms of another element
4. Atoms of different elements combine in simple, whole number ratios to form compounds

   Development of Atomic Theory

   - the concept of an indivisible atom proposed by Dalton inspired other scientists
   - led to the discoveries of sub-atomic particles
     - J.J. Thompson: Z/m of the electron
     - Millikan: Z of the electron
     - Rutherford: the atomic nucleus

Atomic Structure : The electron

J.J. Thompson, 1898 – 1903 :

- the existence of a subatomic particle with a negative charge was demonstrated using a cathode-ray tube: the electron

R. Millikan, 1909 :

- conceived an ingenious experiment to determine the exact charge of the electron
- his setup measured the speed of the fall of tiny droplets of oil
- with this experiment, he established the charge on a single electron
- using the Z/m ratio found by Thompson, the electron's mass was therefore calculated:

$$\text{mass} = Z \bullet m/z$$

$$\text{charge} = -1.60\text{x}10^{-19}\text{C} \bullet \frac{\text{g}}{-1.76\times108\text{C g}}$$

$$= 9.09\times10{-}28 \text{ g}$$

Atomic Structure

E. Rutherford, 1911 (in Montreal) :

- in his experiment, a very thin foil sheet of gold was bombarded with α particles

Modern View of Atomic Structure

The Nucleus:

- Rutherford and Chadwick later discovered that the nucleus contains protons and neutrons
- nuclear radius is 20,000 times smaller than the atomic radius!
- The mass of the atom is due to the nucleus, the volume is due to the orbits of the electrons

Isotopes, Atomic Numbers, and Mass Numbers

- atomic number, Z = number of protons
- atomic mass, A = sum of protons and neutrons
- For the element X, we write:

$${}^{A}_{Z}X$$

1. each element has a unique value of Z
2. isotopes of an element will have the same value of Z, but different values of A (due to number of neutrons)

Isotopes

- the different isotopes of an element usually display similar chemistry, since the number of neutrons has little influence on the reaction's outcome
- we use the atomic mass to identify isotopes:

  $^{235}_{92}U$ is uranium-235 $^{238}_{92}U$ is uranium-238

- the isotopes of hydrogen have special names:

  $^{1}_{1}H$ - hydrogen

  $^{2}_{1}H$ - deuterium

  $^{3}_{1}H$ - tritium

Atomic Masses

- In the SI system, the standard for atomic mass is pure carbon-12:

  Mass of one 12C atom = 12 amu

  where 12 amu is an exact value (i.e. = 12.0000000000000...)

  1 amu (« u ») = 1/12 the mass of a C atom

  = 1.66054 x 10-24 g or 1 g

  = 6.02214 x 1023 u

- The masses of *all other elements* are measured relative to the 12C atom

  Average Atomic Masses

  - IMPORTANT: in nature, the elements exist as a mixture of isotopes
  - we must account for the relative proportions of each isotope in our calculations
  - Atomic mass = weighted average of the respective atomic masses of the natural isotopes of a given element

    $AM = M_1f_1 + M_2f_2 + M_3f_3$…..

    where M = mass of each isotope

    f = fractional abundance of that isotope

  Ex. carbon has two naturally occurring isotopes:

  ^{12}C = 12u : 98.892% ^{13}C = 13.00335u : 1.108%

  The average atomic mass of carbon is:

  (0.98892)(12) + (0.0108)(13.00335) = 12.011 u

Ex. The atomic masses of the two stable isotopes of boron, boron-10 (19.78%) and boron-11 (80.22%), are 10.0129 u and 11.0093 u, respectively. Calculate the average atomic mass of boron.

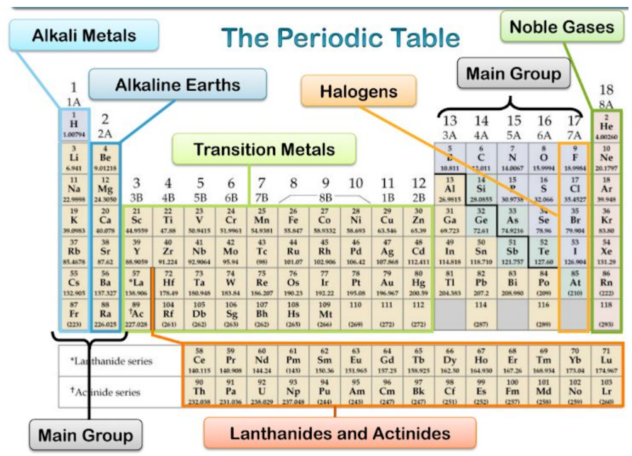
atomic mass of B = $(M_1 \cdot f_1) + (M_2 \cdot f_2)$

$= (0.1978)(10.0129\ u) + (0.8022)(11.0093\ u) = 10.81\ u$

The Periodic Table

D. Mendeleev (1834 – 1907) :

- table in which the elements are grouped together according to their similar physical and chemical properties
- the horizontal rows are called periods
- the vertical columns are called groups (or families)
- three categories of elements:
    - metals (good conductors)
    - non-metals (poor conductors)
    - metalloids or semi-metals



Introduction to Bonding

- electrons are shared : covalent bond nonmetal + nonmetal
- electrons are transferred : ionic bond metal + nonmetal

Covalent Bonds

- atoms bonded covalently form individual, discrete assemblies called molecules

Naming Binary Covalent Compounds :

- non-metal + non-metal
- some compounds have common names:

H_2O = water NH_3 = ammonia

- name lower group number element first
- suffix –ide
- Greek prefixes are used to indicate the numbers of each element in the compound

Name	Prefix	Example
1	mono	NO - nitrogen monoxide
2	di	NO_2 - nitrogen dioxide
3	tri	N_2O_3 - dinitrogen trioxide
4	tetra	N_2O_4 - dinitrogen tetroxide
5	penta	N_2O_5 - dinitrogen pentoxide
6	hexa	SF_6 - sulfur hexafluoride

Ionic Compounds

- Metal + nonmetal
- Metals lose electrons to make positively charged ions = cations
- Non metal gain electrons to make negatively charged ions = anions
- atoms tend to gain/lose electrons to form ions with the same number of electrons as that of the nearest noble gas
- there are no distinct molecules
- ions combine to form an ionic lattice
- NaCl = an ionic compound formed of two ions with charges of +1 and -1
- however, other ionic compounds can have different lattice structures with different charges, but:
- the overall crystal must be electrically neutral

Ex. $Al^{3+} + O^{2-} = Al_2O3$

Naming Binary Ionic Compounds

- name the cation first, then the anion
- prefixes are avoided when possible
- if necessary, the oxidation state of the cation is added as a Roman numeral in brackets

Ex. KBr : potassium bromide

$CaCl_2$: calcium chloride

Al_2O3 : aluminum oxide

FeI_2 : iron (II) iodide

Element	Ion Formula	Systematic Name	Common Name
Chromium	Cr^{2+}	Chromium (II)	Chromous
	Cr^{3+}	Chromium (III)	Chromic
Cobalt	Co^{2+}	Cobalt (II)	
	Co^{3+}	Cobalt (III)	
Copper	Cu^{+}	Copper (I)	Cuprous
	Cu^{2+}	Copper (II)	Cupric
Iron	Fe^{2+}	Iron (II)	Ferrous
	Fe^{3+}	Iron (III)	Ferric
Lead	Pb^{2+}	Lead (II)	
	Pb^{4+}	Lead (IV)	
Mercury	Hg_2^{2+}	Mercury (I)	Mercurous
	Hg^{2+}	Mercury (II)	Mercuric
Tin	Sn^{2+}	Tin (II)	Stannous
	Sn^{4+}	Tin (IV)	Stannic

Sometimes, the anion or cation is polyatomic:

- NH_4^+ : ammonium
- CO_3^{2-} : carbonate
- OH^- : hydroxide

polyatomic ions stay together as a charged unit

Formula	Name
Cations	
NH_4^+	Ammonia

H_3O^+	Hydronium
Hg_2^{2+}	Mercury (I)
Anions	
CH_3COO^- (or $C_2H_3O_2^-$)	Ethanoate (or acetate)
CN^-	Cyanide
OH^-	Hydroxide
ClO^-	Hypochlorite
ClO_2^-	Chlorite
ClO_3^-	Chlorate
ClO_4^-	Perchlorate
NO_2^-	Nitrite
NO_3^-	Nitrate
MnO_4^-	Permanganate
CO_3^{2-}	Carbonate
HCO_3^-	Hydrogen Carbonate (or bicarbonate)
CrO_4^{2-}	Chromate
$Cr_2O_7^{2-}$	Dichromate
O_2^{2-}	Peroxide
PO_4^{3-}	Phosphate
HPO_4^{2-}	Hydrogen Phosphate
$H_2PO_4^-$	Dihydrogen Phosphate
SO_3^{2-}	Sulfite
SO_4^{2-}	Sulfate
HSO_4^-	Hydrogen sulfate (or bisulfate)

Naming Inorganic Compounds

Ex. KNO_3 : potassium nitrate

NH_4Cl : ammonium chloride

$KMnO_4$: potassium permanganate

$Ca(CN)_s$: calcium cyanide

Hydrates :

- an ionic compound containing a fixed number of molecules of water

  $CaCl_s \cdot 6H_2O$ = calcium chloride hexahydrate

  $MgSO_4 \cdot 2H_2O$ = magnesium sulfate dihydrate

Oxoanions

- polyatomic ions of the general formula XO_n^{m-}
- the suffixes –ite and –ate and the prefixes hypo– and per– are used to indicate the oxidation state of the nonmetal atom

  Ex. SO_4^{2-} : sulfate

  SO_3^{2-} : sulfite

  ClO_4^- : perchlorate

  ClO_3^- : chlorate

  ClO_2^- : chlorite

  ClO^- : hypochlorite

Naming Acids

- Brønsted-Lowry definition: acids are proton donors
- *Binary acids (or hydroacids)*
  - a hydroacid is a compound with the general formula H_nX that ionizes in water
  - they are named using the prefix hydro- and the suffix –ic:

    HF : hydrofluoric acid

    HI : hydroiodic acid

    H_2S : hydrosulfuric acid

Naming oxoacids

- an oxoacid is a compound with the general formula H_mXO_n that ionizes in water
- oxoacid = oxoanion + H+ ions
- their names are based on the oxoanion from which they are formed :

ite becomes –ous

ate becomes –ic

Ex. ClO^- (hypochlorite) → hypochlorous acid

ClO_2^- (chlorite) → chlorous acid

ClO_3^- (chlorate) → chloric acid

ClO_4^- (perchlorate) → perchloric acid

Masses from Chemical Formula

Covalent compounds :

- Molecular mass = sum of the atomic masses of each atom in the molecule (in u)

Ionic compounds :

- Formula mass = sum of the atomic masses of each atom in the formula unit (in u)

Ex. What is the molecular mass of H_2O

$$2 \text{ atoms of H} \times \frac{1.008 \text{ u}}{1 \text{ H atom}} = 2.016 \text{ u}$$

$$1 \text{ atom O} \times \frac{15.999 \text{ u}}{1 \text{ O atom}} = 15.999 \text{ u}$$

1 Molecule of H_2O = 2.016 u + 15.999 u = 18.015 u

Molecule Mass vs. Exact Mass

Glucose molecular formula : $C_6H_{12}O_6$

Molecular Mass: Use the average atomic mass of each element

6 x 12.01 + 12 x 1.01 + 6 x 16.00 = 180.18 u

Exact Mass: Use exact mass of the most abundant isotope of each element

6 x 12.000000 + 12 x 1.007825 + 6 x 15.994915 = 180.06339 u