

# CHEM 1211 - Principles of Chemistry I

## Chapter 2 - Atoms, Molecules, and Ions

### 2.1 The Early History of Chemistry

- chemistry has been important since ancient times
- the processing of nature's ores to produce metals for ornaments and weapons and the use of embalming fluids are two applications of chemical phenomena that were utilized prior to 1000 B.C.
- the Greeks were the first to try explain why chemical changes occur; by about 400 B.C. they had proposed that all matter was composed of four fundamental substances: fire, earth, air, and water
  - they also proposed that matter was made up of small indivisible particles; supporters of this view were **Demokritos of Abdera** (c. 460-c. 370 B.C) and Leucippos, who used the term *atomos* (which later became *atoms*)
  - Greeks had no experiments to test their ideas
- next two thousand years dominated by *alchemy* (a pseudoscience)
  - alchemists often mystics and fakes but did discover several elements and learned to prepare mineral acids
- modern chemistry began in the sixteenth century with the development of systematic metallurgy (extraction of metals from ores) by German chemist **Georg Bauer** (1494-1541), and the medicinal application of minerals by a Swiss alchemist/physician known as **Paracelsus** (full name: Phillipus Theophrastus Bombastus von Hohenheim)
- the first "chemist" to perform quantitative experiments was **Robert Boyle** (1627-1691), who carefully measured the relationship between the pressure and volume of air
  - when Boyle published *The Sceptical Chymist* in 1661, the quantitative sciences of physics and chemistry were born
- the phenomenon of combustion evoked intense interest in the seventeenth and eighteenth centuries
  - the German chemist **Georg Stahl** (1660-1734) suggested a substance he called "phlogiston" flowed out of the burning material; a substance in a closed container would go out because the vessel became filled with phlogiston
  - oxygen gas, discovered by **Joseph Priestly** (1733-1804), found to support vigorous combustion and was supposed to be low in phlogiston; oxygen was originally called "dephlogisticated air"

### 2.2 Fundamental Chemical Laws

- by the late eighteenth century, combustion had been studied extensively; carbon dioxide, nitrogen, hydrogen, and oxygen had been discovered; and the list of elements continued to grow
- it was the French chemist **Antoine Lavoisier** (1743-1794) who finally explained the true

nature of combustion by measuring the weights of the reactants and products and determined that *mass is neither created nor destroyed*

- Lavoisier's verification of this **law of conservation of mass** was the basis for the developments in chemistry in the nineteenth century
  - in 1789 Lavoisier published the first modern chemistry textbook, *Elementary Treatise on Chemistry*, in which he presented a unified picture of the chemical knowledge assembled up to that time
- Frenchman **Joseph Proust** (1754-1826) showed that *a given compound always contains exactly the proportion of elements by mass*
  - the principle of constant composition of compounds, originally called "Proust's law" is now known as the **law of definite proportion**
- Proust's discovery stimulated **John Dalton** (1766-1844), an English schoolteacher, to think about atoms as the particles that might compose elements; a given compound should always contain the same combination of these atoms
  - Dalton discovered another principle that convinced him even more of the existence of atoms; he noted that carbon and oxygen form two different compounds that contain different relative amounts of carbon and oxygen
  - the principle is known as the **law of multiple proportions**: *when two elements form a series of compounds, the ratios of the masses of the second element that combine with 1 gram of the first element can always be reduced to small whole numbers*

### 2.3 Dalton's Atomic Theory

- in 1808 Dalton published *A New System of Chemical Philosophy*, in which he presented his theory of atoms (see the four points listed on page 43)
- because the mass ratio of oxygen:hydrogen was 8:1, Dalton thought the atomic mass of oxygen was 8 and that of hydrogen 1 (nature would be as simple as possible)
- the keys to determining absolute formulas for compounds were provided in the experimental work of the French chemist **Joseph Gay-Lussac** (1778-1850) and by the hypothesis of Italian chemist **Amadeo Avogadro** (1776-1856)
  - in 1809 Gay-Lussac performed an experiment - representation shown in Figure 2.4
  - in 1811 Avogadro interpreted these results by proposing that *at the same temperature and pressure, equal volumes of different gases contain the same number of particles* (known as **Avogadro's hypothesis**)
    - 2 molecules of hydrogen react with 1 molecule of oxygen to give 2 molecules of water

### 2.4 Early Experiments to Characterize the Atom

- now scientists began to wonder about the nature of the atom;
  - (1) what is an atom made of, and
  - (2) how do the atoms of the various elements differ

## *The Electron*

- English physicist, **J. J. Thomson** (1856-1940), studied electrical discharges in partially evacuated tubes called **cathode ray tubes** during the period 1898-1903
  - he found that when high voltage was applied to the tube, a "ray" he called a **cathode ray** (because it emanated from the negative electrode, or cathode) was produced
  - Thomson postulated that the ray was a stream of negatively charged particles, now called **electrons**
  - he determined the *charge-to-mass ratio* of an electron
  - electrons were produced from electrodes made of different metals - speculated all atoms contain electrons
  - since he found negatively charged particles also speculated that atoms would have positive charged particles
  - see Figure 2.9; Thomson's plum pudding model of the atom
- in 1909 **Robert Millikan** (1868-1953) working at the U. of Chicago was able to determine the charge on an electron and therefore able to calculate the mass of an electron
  - see Fig. 2.10

## *Radioactivity*

- in 1896 **Henri Becquerel** discovered what he called **radioactivity**, the spontaneous emission of radiation from uranium
- in the early twentieth century alpha, beta, and gamma radiation were discovered

## *The Nuclear Atom*

- in 1911 **Ernest Rutherford** (1871-1937) carried out an experiment to test Thomson's plum pudding model of the atom
  - see Figure 2.12
  - see Figure 2.13
- Rutherford's experiment's gave credence to the idea of a **nuclear atom** - an atom with a dense center of positive charge (the **nucleus**) with electrons moving around the nucleus at a distance that is large relative to the nuclear radius

## **2.5 The Modern View of Atomic Structure: An Introduction**

- nucleus contains protons and neutrons; electrons surround the nucleus
  - see Table 2.1
- the number of protons determines the element
- **isotopes** are atoms that differ in the number of neutrons but have the same number of protons
- symbol for an atom is  ${}^A_Z\text{X}$ , where **Z** is the **atomic number** (number of protons) and **A** is the **mass number** the total number of protons and neutrons

- number of neutrons =  $A - Z$
- $Z$  determines the element

## 2.6 Molecules and Ions

- for that hold atoms together in compounds are called **chemical bonds**
  - have **covalent bonds** (sharing of electrons) and **ionic bonds** (electrostatic attraction between ions)
    - water, ammonia, and methane are examples of covalent compounds
    - sodium chloride, calcium bromide, lead nitrate are examples of ionic compounds (made up of cations and anions)
- the ammonium ion and nitrate ion are examples of **polyatomic ions**

## 2.7 An Introduction to the Periodic Table

- a **periodic table** is the most valuable tool a chemist has
- organization is based on atomic number
- names come from various sources
  - see Table 2.2
- most elements are **metals**; characteristics of metals are that they
  - can conduct heat
  - can conduct electricity
  - are malleable
  - are ductile, and
  - have a lustrous appearance
    - see Fig. 2.21
  - metals tend to lose electrons, i.e., for positive ions
- there are relatively few **nonmetals**; characteristics of nonmetals are the opposite of metals
- periodic table arranged so that elements that have similar properties are grouped together
  - alkali metals, alkaline earth metals, transition metals, halogens, noble gases, lanthanides, actinides
- vertical **groups** and horizontal **periods** in the periodic table

## 2.8 Naming Simple Compounds

- in the early days of chemistry no systematic method for naming compounds, e.g., sugar of lead, blue vitrol, quicklime, Epsom salts, milk of magnesia, gypsum, and laughing gas
  - as more compounds were identified a system was needed

### *Binary Ionic Compounds (Type 1)*

- see Table 2.3

### ***Binary Ionic Compounds (Type 2)***

- see Table 2.4
  - note that Iron (II) is also known as ferrous ion and Iron (III) is known as ferric ion, Copper (I) - cuprous ion and Copper (II) - cupric
  - elements that form only one cation do not need to be identified by a Roman numeral (e.g., silver ion)

### ***Ionic Compounds with Polyatomic Ions***

- see Table 2.5 page 65

### ***Binary Compounds (Type III; Covalent - Contain Two Nonmetals)***

- see Table 2.6

### ***Formulas from Names***

- see sample exercise

### ***Acids***

- when dissolved in water, certain molecules produce a solution containing free  $H^+$ 
  - see Table 2.7
  - see Table 2.8
    - note anion of nitrous acid is nitrite ion and of nitric acid, nitrate ion