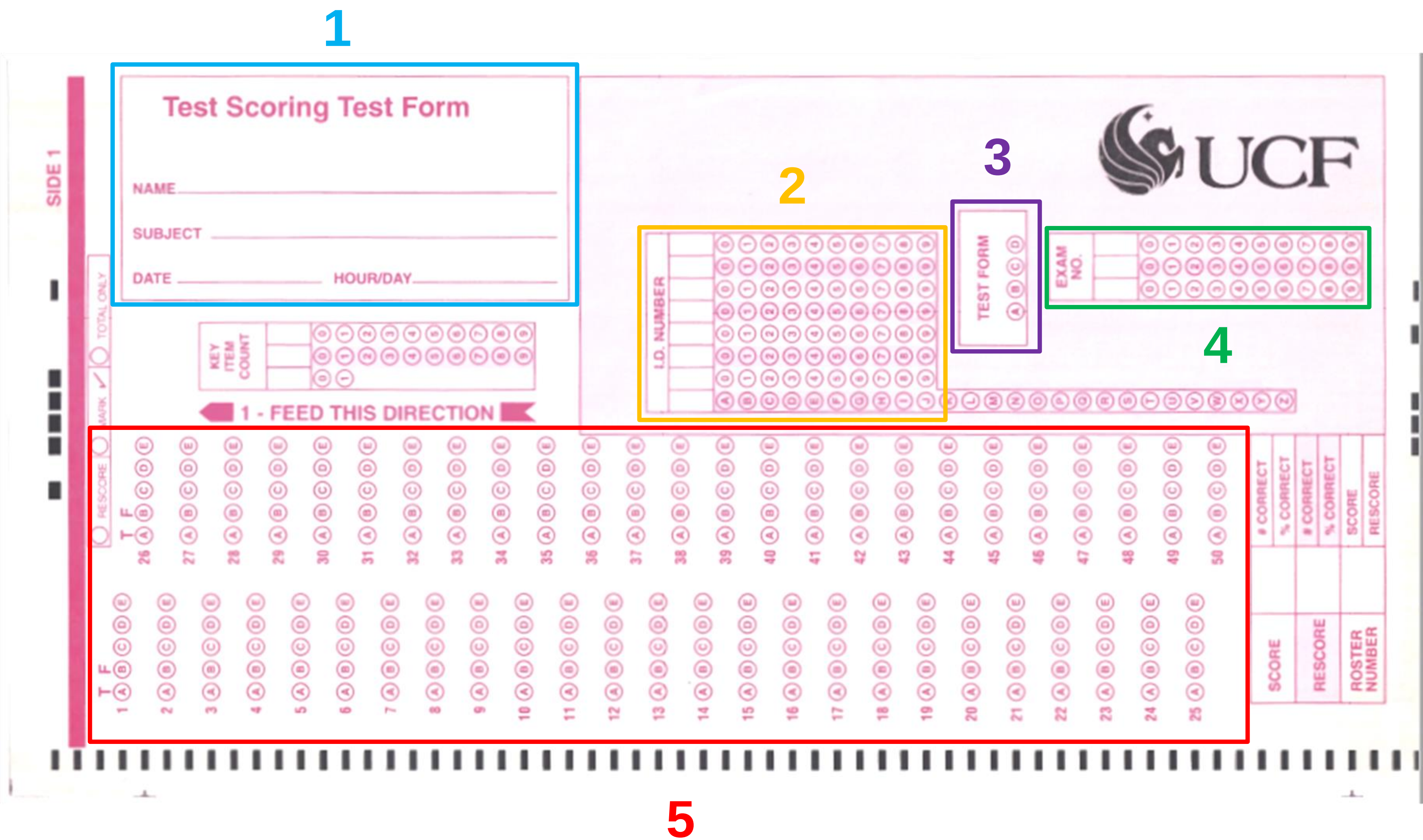
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Chapter 1

Introduction to Chemistry

Gang Chen, Department of Chemistry, UCF

The Macroscopic Perspective

Matter is anything that has mass and can be observed.

Matter is observed through two types of changes.

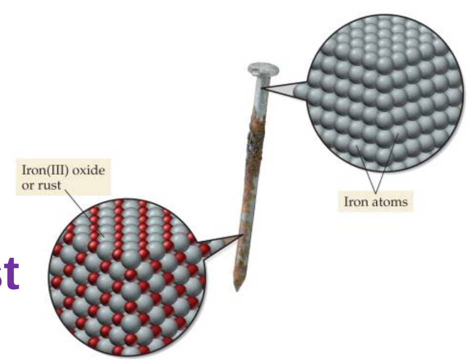
i) Physical changes

melting & freezing

boiling & condensing

dissolving

ii) Chemical changes



iron

rust

Practice 1.8, 1.13

Example Problem

Q. Classify each <u>change</u> as physical or chemical!

a) rushing of an iron bridge **C**

b) melting of ice **P**

c) burning of a wooden stick **C**

d) dissolving of sugar in water **P**

e) digestion of a baked potato **C**

Something new is formed in a chemical change!

The Macroscopic Perspective

- Physical properties are variables of matter that we can measure without changing the identity of the substance being observed.

Color	Luster	Solubility
Size	Hardness	Conductivity
Odor	Condensing	Boiling point
Density	Melting point	

More physical properties can be found online
https://en.wikipedia.org/wiki/Physical_property

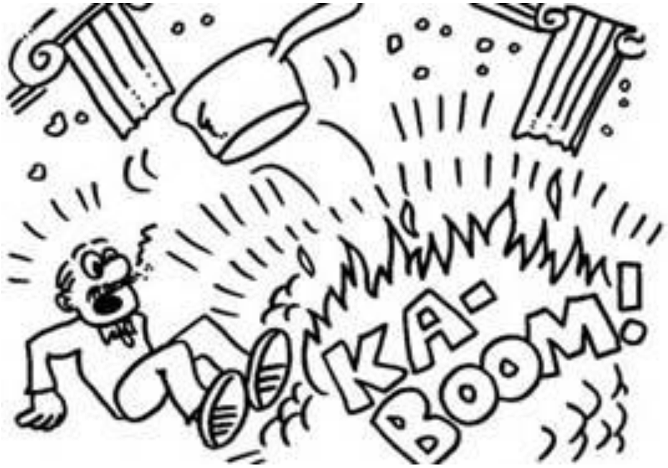
The Macroscopic Perspective

- Chemical properties are determined only by observing how a substance changes its identity in chemical reactions.

oxidizing-reducing
(corrosion)

flammability

explosiveness



Practice 1.16

Example Problem

Q. Which of the following properties of a metal are chemical properties?

a) It is hard **P**

b) It rusts in air **C**

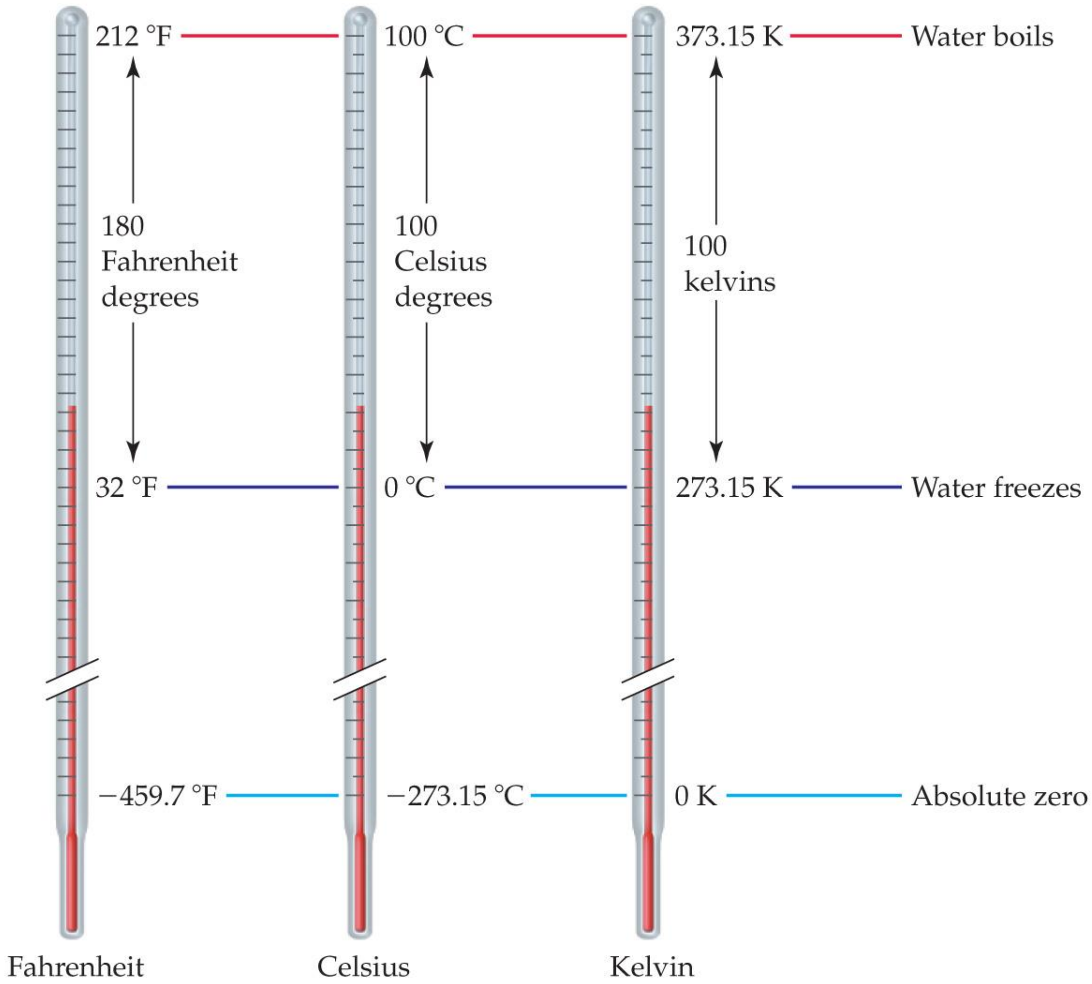
c) Its density is 5.5 g/cm3 **P**

d) It reacts with a base **C**

Numbers and Measurements in Chemistry

- Chemists quantify data, express collected data with units and significant figures.
  - Units - designate the type of quantity measured.
  - Prefixes - provide scale to a base unit.
  - Significant Figures - indicate the amount of information that is reliable when discussing a measurement.

Temperature



Temperature is measured using the Fahrenheit, Celsius, and Kelvin (absolute) temperature scales.

Temperature Scale Conversions

$$^{o}F = (1.8 \times {}^{o}C) + 32$$

Q: -320.4 ^{o}F in K

$$\boxed{^{o}C = (^{o}F - 32)/1.8}$$

^{o}C = (-320.4 - 32) / 1.8
= (-352.4) / 1.8
= -195.8

$$\boxed{K = {}^{o}C + 273.15}$$

K = -195.8 + 273.15 = 77.4

$$^{o}C = K - 273.15$$

Practice 1.41, 1.42

Chapter 2

Atoms and Molecules

Gang Chen, Department of Chemistry, UCF

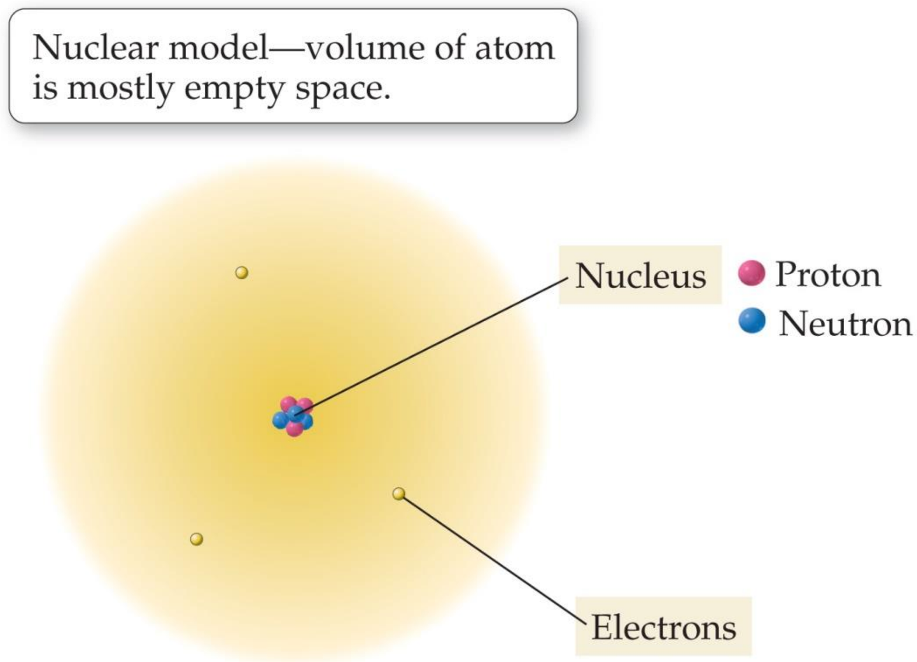
Atomic Structure and Mass

Atoms have a <u>nucleus</u> which contains <u>protons</u> and <u>neutrons</u>.
The nucleus is surrounded by a cloud of <u>electrons</u>.

Most of the atom’s mass (proton; neutron) & its positive charge (proton) are in the nucleus.

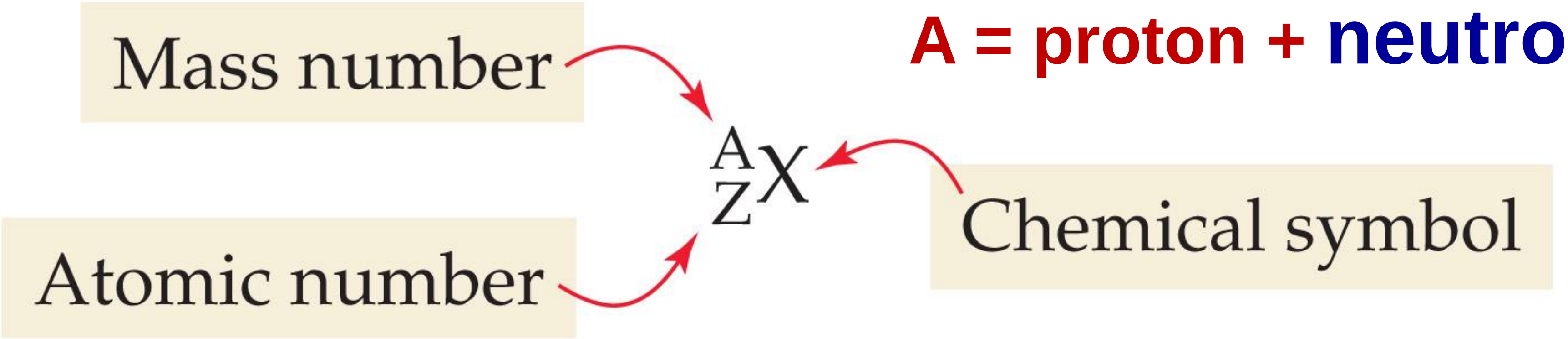
The number of negatively charged electrons = number of positively charged protons.

Therefore, the <u>atom</u> is electrically neutral.



Atomic Number and Mass Number

- Atomic Number, *Z*, is the number of protons in a nucleus.
  - identifies the element
- Mass Number, *A*, is the sum of the number of protons and number of neutrons in a nucleus.



A = proton + neutron

Practice 2.9, 2.10

Example Problem

Q. How many protons, neutrons, and electrons are in the ^{22}Na atom?

Z = 11, A = 22, so A – Z = 11

a) 11 protons, 11 neutrons, 10 electrons

b) 22 protons, 11 neutrons, 11 electrons

c) 11 protons, 11 neutrons, 11 electrons

d) 10 protons, 12 neutrons, 10 electrons

e) 11 protons, 22 neutrons, 11 electrons

Isotopes

- Isotopes are atoms of an element that differ in the number of neutrons in their nucleus.

  - same *Z* but different *A*

E.g., the symbols for the isotopes of carbon are:

$$^{12}_{6}\mathrm{C} \qquad ^{13}_{6}\mathrm{C} \qquad ^{14}_{6}\mathrm{C}$$

C-12 **C-13** **C-14**

- Isotopic abundance is the mass percentage of an isotope in a naturally occurring element.

Practice 2.12, 2.13, 2.14

Atomic Masses

- Relative atomic mass for an element is an *average* of the atomic masses for the naturally occurring isotopes for an element.

  - Carbon-12 = 12.0000 x 0.9893 = 11.87 amu
  - Carbon-13 = 13.0036 x 0.0107 = 0.139 amu

  - Average mass = 11.87 + 0.139 = 12.01 amu

Practice 2.19, 2.21

Ions

Atoms acquire charge (form ions) by gaining or losing <u>electrons</u> (not protons!!!) in chemical reactions to form ions.

Atoms:

Al: 13p^+, 13e^- **O: 8p^+, 8e^-** **Ca: 20p^+, 20e^-**

Ions:

Al^{3+}: 13p^+,10e^- **O^{2-}: 8p^+, 10e^-** **Ca^{2+}: 20p^+, 18e^-**

monatomic cation **monatomic anion** **monatomic cation**

polyatomic cation **NH_4^+** **polyatomic anion** **NO_3^-**

Practice 2.26, 2.27

Chemical Formulas

- Compound: a pure substance; made up of atoms of two or more elements (CO; NO;…..).

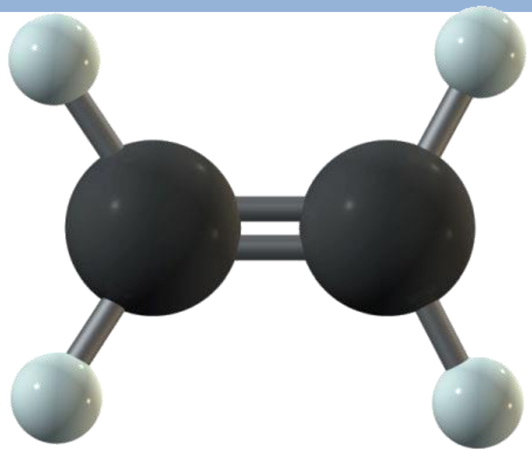
- Chemical formulas describe a compound in terms of the elements the compound contains.
  - The number of atoms for each element is indicated by a subscript to the right of the chemical symbol.
    - Groups of atoms can be designated using parentheses. Subscripts outside these parentheses mean that all atoms enclosed in the parentheses are multiplied by the value indicated by the subscript.

$$Fe_3(PO_4)_2 \bullet 8H_2O$$

3 Fe, 2 P, 16 O, 16 H

Practice 2.35, 2.36

Chemical Formulas



Ethylene, C_2H_4

- Molecular formulas indicate the elements and number of atoms of each element actually contained in a discrete unit of a compound.

$$C_2H_4$$

- Empirical formulas indicate the smallest whole number ratio between the number of atoms of each element in a molecular formula.

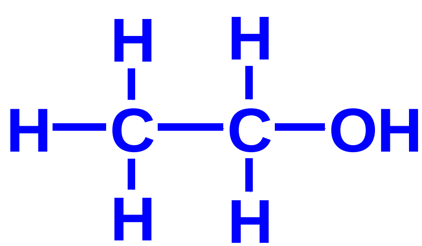
$$CH_2 \qquad (CH_2)_n$$

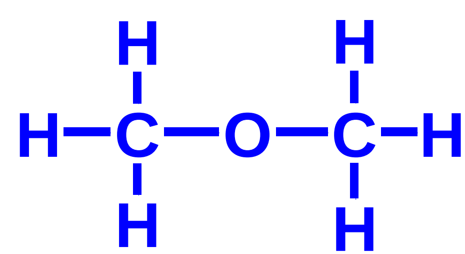
Practice 2.47

Inorganic and Organic Chemistry

- Organic chemistry is the study of the compounds of the element carbon, usually with oxygen, nitrogen, and hydrogen.
    - More than 18 million organic compounds exist.
    - Includes biological molecules and nearly all synthetic polymers.
    - Isomers: Different organic molecules that have the same formula but are connected differently.

C_2H_6O

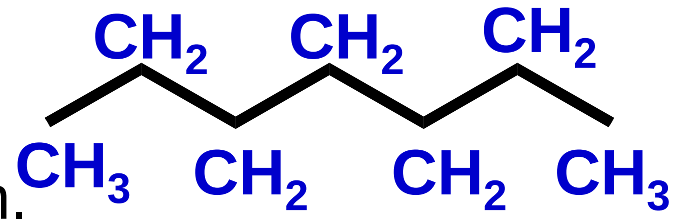




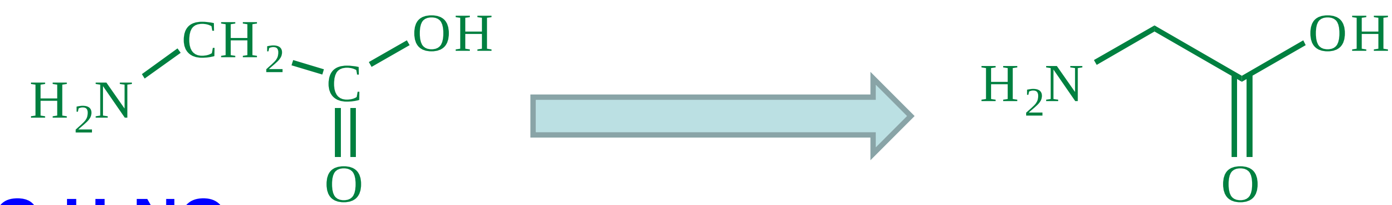
- Inorganic chemistry is the study of all other elements and their compounds.

Organic Chemistry

- Because carbon compounds can become quite large, organic compounds are described simply and unambiguously using line structures, where carbons and hydrogens are not explicitly shown.



  - Each corner or end of a line is a carbon.
  - Hydrogen atoms on carbon atoms are implied. Carbon makes four bonds, “missing” bonds go to hydrogen atoms. Hydrogen can only make one covalent bond to another atom.
  - Hydrogen atoms on any other element are shown
  - All other elements are shown

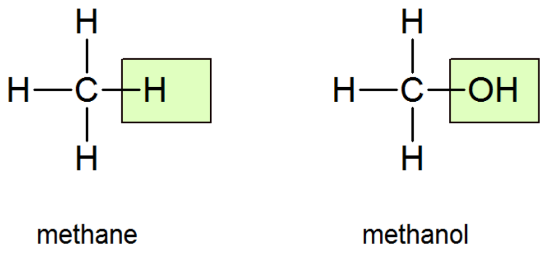


$C_2H_5NO_2$

Practice 2.66-2.70

Functional Groups

- Functional groups are arrangements of atoms that tend to display similar chemical properties.
    - Chemical formulas are often written to emphasize functional groups.
        - Methanol, an alcohol, is often written CH_3OH instead of CH_4O.
- Hydrocarbons contain only H and C atoms.



- Addition of functional groups to hydrocarbons results in more complex compounds.

Practice 2.65

Chemical Nomenclature

- Chemical nomenclature is a systematic means of assigning names to chemical compounds.

- Binary compounds contain <u>only two elements</u>.
  - Covalent binary compounds are named differently from ionic binary compounds.
  - Recognizing a compound as ionic or covalent assists in naming.
    - A metal and a nonmetal generally combine to form ionic compounds.
    - Two nonmetals generally combine to form a covalent compound.
    - Presence of polyatomic ions indicates ionic bonding.

Practice 2.71, 2.72

Naming Covalent Compounds

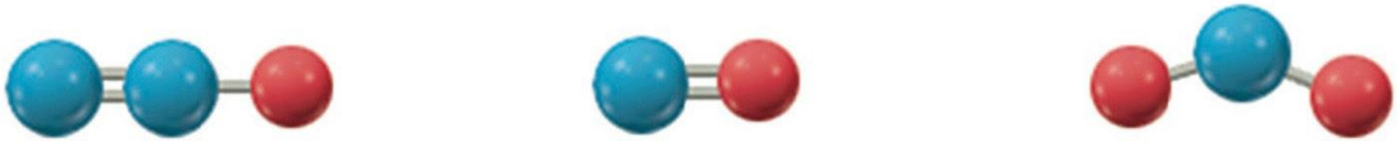
- The first element in the formula retains is full name.
- The second element is named by replacing the ending from its name with the suffix -ide.
- Both elements are preceded by a number-designating prefix except when there is only one atom of the first element, which will not use the prefix mono-.

Table 2.4

Greek prefixes for the first ten numbers

Number	Prefix
One	Mono-
Two	Di-
Three	Tri-
Four	Tetra-
Five	Penta-
Six	Hexa-
Seven	Hepta-
Eight	Octa-
Nine	Nona-
Ten	Deca-

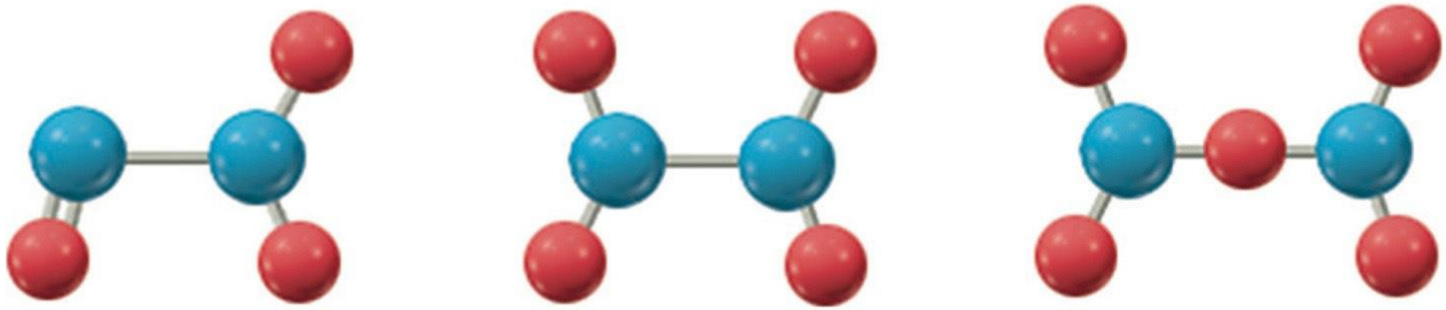
Naming Covalent Compounds



(a) Dinitrogen monoxide, N_2O

(b) Nitrogen monoxide, NO

(c) Nitrogen dioxide, NO_2



(d) Dinitrogen trioxide, N_2O_3

(e) Dinitrogen tetroxide, N_2O_4

(f) Dinitrogen pentoxide, N_2O_5

Nitrogen forms a number of binary compounds with oxygen.

Practice 2.73-2.75

Naming Ionic Compounds

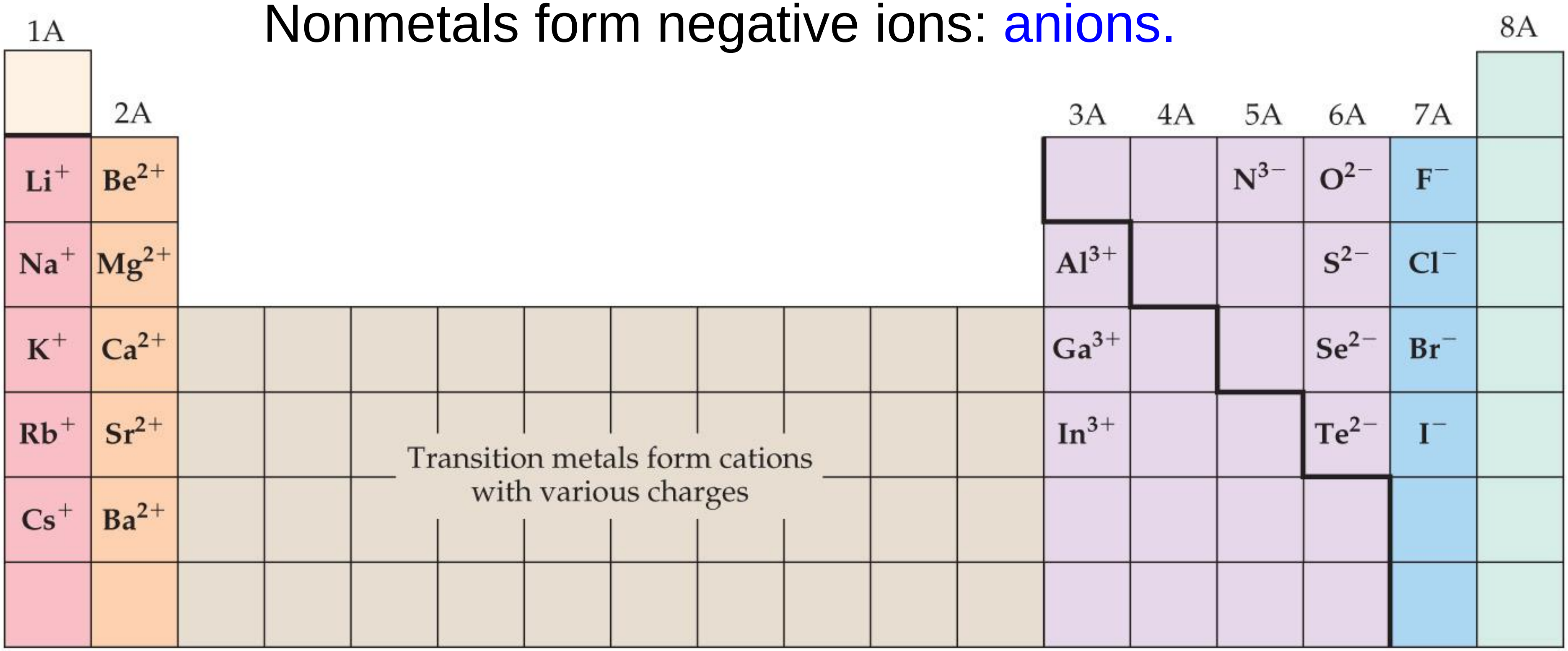
- Ionic compounds are electrically neutral and are named in order of “cation anion”, as in sodium chloride.
  - The cation retains its full name.
    - Monoatomic cation charge can often be found by position in the periodic table.
    - Cations with more than one charge (e.g., transition metals) are named using Roman numerals indicating the charge, e.g., iron(II)
  - Monatomic anions are named by replacing the ending of the element name with the suffix -ide, e.g., brom*ide*
  - A polyatomic cation or anion is named using its common name.

copper(I) oxide	Cu_2O
iron(II) chloride	$FeCl_2$

Naming Ionic Compounds

Ion charge (some) predicted by group number on periodic table:

Metals form positive ions: cations.
Nonmetals form negative ions: anions.



Metal: charge on cation = grp number

Nonmetal: charge on anion = grp number - 8

Example Problem

Q: The correct formula for potassium phosphate is:

a) KPO_4 b) K_2PO_4 c) $\boxed{K_3PO_4}$ d) $K(PO_4)_2$ e) $K(PO_4)_3$

K^+ (1+) x 3 = 3+

PO_4^{3-} (3-) x 1 = 3-

Practice 2.76-2.79

Chapter 3

Molecules, Moles, and Chemical Equations

Gang Chen, Department of Chemistry, UCF

Balancing Chemical Equations

- The law of conservation of matter: matter is neither created nor destroyed.
  - Chemical reactions must obey the law of conservation of matter.
    - The same number of atoms for each element must occur on both sides of the chemical equation.
    - A chemical reaction simply rearranges the atoms into new compounds.

Q. Write a balanced chemical equation describing the reaction between butane (C_4H_{10}) and oxygen (O_2) to form carbon dioxide and water.

$$C_4H_{10(g)} + O_{2\,(g)} \rightarrow CO_{2\,(g)} + H_2O_{(g)}$$

balance C

$$C_4H_{10(g)} + O_{2\,(g)} \rightarrow 4CO_{2\,(g)} + H_2O_{(g)}$$

balance H

$$C_4H_{10(g)} + O_{2\,(g)} \rightarrow 4CO_{2\,(g)} + 5H_2O_{(g)}$$

balance O

$$C_4H_{10(g)} + 6.5O_{2\,(g)} \rightarrow 4CO_{2\,(g)} + 5H_2O_{(g)}$$

$$2C_4H_{10(g)} + 13O_{2\,(g)} \rightarrow 8CO_{2\,(g)} + 10H_2O_{(g)}$$

Practice 3.11, 3.13, 3.17

Chemical Equations for Aqueous Reactions

- Molecular equation

$$AgNO_3(aq) + NaCl(aq) \longrightarrow AgCl(s) + NaNO3(aq)$$

- Total ionic equation

$$Ag^+(aq) + NO_3^-(aq) + Na^+(aq) + Cl^-(aq) \longrightarrow$$
$$AgCl(s) + NO_3^-(aq) + Na^+(aq)$$

- Net ionic equation

$$Ag^+(aq) + Cl^-(aq) \longrightarrow AgCl(s)$$

Spectator ions: NO_3^- and Na^+

Spectator ions are ions uninvolved in the chemical reaction.

Acid-Base Reactions

- Acids are substances that dissolve in water to produce H^+ (or H_3O^+) ions.
  - Examples: HCl, HNO_3, H_3PO_4, HCN

- Bases are substances that dissolve in water to produce OH^- ions.
  - Examples: NaOH, $Ca(OH)_2$, NH_3

$$HCl(g) + H_2O(l) \longrightarrow H_3O^+(aq) + Cl^-(aq)$$

$$NaOH(s) \longrightarrow Na^+(aq) + OH^-(aq)$$

Acid-Base Reactions

- Mixing an acid and a base leads to a reaction known as neutralization, in which the resulting solution is neither acidic nor basic.
  - Net ionic equation for neutralization of strong acid and strong base.

salt

$$2HNO_{3(aq)} + Ba(OH)_{2(aq)} \rightarrow Ba(NO_3)_{2(aq)} + 2H_2O(l)$$

$$2H^+_{(aq)} + \cancel{2NO_3^-}_{(aq)} + \cancel{Ba^{2+}}_{(aq)} + 2OH^-_{(aq)} \rightarrow$$

$$\cancel{Ba^{2+}}_{(aq)} + \cancel{2NO_3^-}_{(aq)} + 2H_2O(l)$$

$$H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O(l)$$

the net ionic equation for strong acid-strong base rxn.

Precipitation Reactions

- A precipitation reaction is an aqueous reaction that produces a solid, called a precipitate.

$$Pb(NO_3)_2(aq) \; + \; 2\,NaI(aq) \;\rightarrow\; PbI_2(s) \; + \; 2\,NaNO_3(aq)$$

- Net ionic reaction for the precipitation of lead(II) iodide.

$$Pb^{2+}(aq) + 2I^{-}(aq) \longrightarrow PbI_2(s)$$

Avogadro's Number and the Mole

- A mole is a means of counting the large number of particles in samples.
  - One mole is the number of atoms in exactly 12 grams of ^{12}C (carbon-12) - standard.
  - 1 mole contains Avogadro's number (6.022×10^{23} particles/mole) of particles.
  - The mass of 6.022×10^{23} atoms of any element is the molar mass of that element.
  - The molar mass of a compound is the sum of the molar masses of ALL the atoms/ions in a compound.

Practice 3.38, 3.39

Calculations Using Moles and Molar Mass

- Avogadro's number functions much like a unit conversion between moles to number of particles.

$$moles\ of\ a\ substance = \frac{number\ of\ particles}{Avogadro's\ number}$$

- Molar mass ($C_6H_{12}O_6$) = 180.15 g/mol

- How many O atoms are present in 214 g of mannose?

$$\textbf{214 g mannose} \times \frac{1\,mole}{180.15\ g\ mannose} \times \frac{6.022\times10^{23}\ molecules}{1\,mole} \times \frac{\textbf{6 O atoms}}{\textbf{1 molecule}}$$

$$\mathbf{4.29\times10^{24}}\ \textbf{O atoms}$$

Practice 3.51, 3.53

Elemental Analysis: Determining Empirical and Molecular Formulas

- Empirical formulas can be determined from an elemental analysis.
  - An elemental analysis measures the mass percentage of each element in a compound.
  - The formula describes the composition in terms of the atomic ratio of each element.
  - The molar masses of the elements provide the connection between the elemental analysis and the formula.

Elemental Analysis: Determining Empirical and Molecular Formulas

- Assume a 100 gram sample size
- Percentage element $\times$ sample size = mass element in compound. (e.g., 16% carbon = 16 g carbon)
- Convert mass of each element to moles using the molar mass.
- Divide by smallest number of moles to get mole to mole ratio for empirical formula.
- When division by smallest number of moles results in small rational fractions, multiply all ratios by an appropriate integer to give whole numbers. Empirical formulas do not have fractions!
    - $2.5 \times 2 = 5$, $1.33 \times 3 = 4$, etc.

Elemental Analysis: Determining Empirical and Molecular Formulas

Q. Determine the simplest formula of the compound which has the composition 74.0 % C, 8.65 % H, 17.4 % N.

$$74.0\ g\ C \times \frac{1\ mol\ C}{12.01\ g\ C}$$ **6.16 mol C**

$$8.65\ g\ H \times \frac{1\ mol\ H}{1.008\ g\ H}$$ **8.58 mol H**

$$17.4\ g\ N \times \frac{1\ mol\ N}{14.01\ g\ N}$$ **1.24 mol N**

$$C_{6.16}H_{8.58}N_{1.24} \rightarrow C_{\frac{6.16}{1.24}}H_{\frac{8.58}{1.24}}N_{\frac{1.24}{1.24}}$$

$$C_{4.97}H_{6.92}N \rightarrow \text{e. f.} = C_5H_7N$$

Practice 3.59, 3.61

Molarity

- Molarity, or molar concentration, *M*, is the number of moles of solute per liter of solution.
  - Provides relationship among molar concentration, moles of solute, and liters of solution.

$$\text{Molarity } (M) = \frac{\text{moles of solute}}{\text{liter of solution}}$$

- If we know any two of these quantities, we can determine the third.

$$M = n / V \qquad n = M \times V$$

Practice 3.63, 3.66

Dilution

- Dilution is the process in which solvent is added to a solution to decrease the concentration of the solution.

  - The number of moles of solute is the same before and after dilution.

  - Since the number of moles of solute equals the product of molarity and volume ($M \times V$), we can write the following equation, where the subscripts denote initial and final values.

$$M_{\mathrm{i}} \times V_{\mathrm{i}} = M_{\mathrm{f}} \times V_{\mathrm{f}}$$

Example Problem

Q. Determine the initial volume needed to generate 10.0 L of 0.45 M solution from a 3.0 M solution

Use $M_i \times V_i = M_f \times V_f$

M_f = 0.45 M; V_f = *10.0 L*

M_i = 3.0 M; V_i = *?*

$$\mathbf{V_i = \frac{M_f \times V_f}{M_i} = \frac{0.45\,M \times 10.0\,L}{3.0\,L}} = \mathbf{\mathit{1.5\ L}}$$

Dilution: adding more solvent to a solution

Practice 3.67, 3.69