

Chapter 2

Atoms, Molecules, and Ions

Chapter 2 suggested problems

10th Ed. - 2.x: 7, 15, 21, 23, 25, 29, 37, 47, 49, 53, 55, 59, 61, 63, 65, 92, 95, 97

11th Ed. - 2.x: 9, 17, 23, 25, 27, 31, 39, 49, 51, 55, 59, 65, 67, 69, 71, 98, 102, 104

Important web sites: **WebElements** (<http://www.webelements.com/>)

Class Notes

- I. The atomic theory of matter and the discovery of atomic structure: a bit of history
 - A. 5th Century B.C. - Leucippus and his student Democritus postulate that all matter is made of invisible, indivisible particles called "atomos" (indivisible)
 1. At roughly the same time Empedocles proposed that everything is made up of four basic elements: earth, air, water, and fire
 2. This notion was later popularized by the work of Aristotle (abt. 200 B.C.)
 - B. 1st Century B.C. - the rejected ideas of Leucippus and Democritus were popularized by the Roman poet Lucretius in "De Rerum Naturae"
 - C. These ideas were rediscovered in the late 16th Century; by the late 17th Century physicists such as Robert Boyle, Robert Hooke, and Isaac Newton began to suggest that the physical properties of matter might be explained by forces that acted upon some form of ultimate building block of matter
 1. Newton in "Optiks": it seems probable to me that God in the beginning formed matter in solid, massy, hard, impenetrable, moveable particles. . . ."
 - D. 1803-1807 - John Dalton and his atomic theory
 1. Elements are made of atoms
 2. All of the atoms in an element are identical
 3. The atoms of an element are different from the atoms in any other element
 4. The atoms of two or more elements can combine to form compounds

5. Law of constant composition: compounds always have the same ratios of atoms regardless of where the compound is found
6. Conservation of matter: atoms are not created or destroyed in chemical reactions

E. mid-1800s - The discovery of cathode-ray tubes

1. 1897 - J. J. Thomson discovers the electron
2. 1909 - Millikan discovers the charge and mass of electrons

F. 1896 - Radioactivity - the divisibility of indivisible particles becomes apparent

1. 1896 - Henri Becquerel discovers the "radioactivity" of a uranium-bearing compound - pitchblende
2. abt. 1898 - Marie and Pierre Curie discover two additional radioactive elements - radium and polonium

G. Ernest Rutherford

1. 1909-10 - the discovery of the nucleus
 - a. The nucleus of the atom contains 99.95% of the mass of the atom but 10^{-15} of its volume
 - b. If the nucleus of an atom is the size of a football stadium, the atom itself would be the size of the earth!
2. 1919 - discovery of the proton, from "protos" the first identified building block of all atomic nuclei
3. 1932 - James Chadwick discovers the neutron

II. The modern view of atomic structure

A. A summary

	charge	relative mass	absolute mass (kg)
proton	+1	1	1.6726×10^{-27}
neutron	0	1	1.6749×10^{-27}
electron	-1	1/1800	9.109×10^{-31}

- B. Atomic radii: about 100-500 pm
- C. The diameters of atomic nuclei: around 0.01 pm or less
- D. Atoms are mostly empty space, just as the Solar System is mostly empty space

III. Atomic numbers and atomic weights

A. Protons

1. The number of protons in the nucleus is unique for each type of atom (i.e., the atoms in each element); the number of protons in an element is invariable - if the number of protons changes, the element itself changes
2. Atomic number - the top number for each element in the Periodic Table
3. Abbreviated "Z"

B. Neutrons

1. Protons and neutrons are held together in the nucleus at the center of the atom
2. Despite mutual repulsion of like-charged particles they are held together by the Strong Force, which is much stronger than the repulsion caused by the electromagnetic force law
3. *Isotopes* - substances with the same number of protons but different numbers of neutrons
4. All elements have one isotope; virtually all naturally occurring elements consist of two or more isotopes
 - a. *Atomic mass* - the sum of the protons and neutrons in an isotope
 - b. $^{12}_6\text{C}$ $^{13}_6\text{C}$ $^{14}_6\text{C}$
 - c. $^{35}_{17}\text{Cl}$ $^{37}_{17}\text{Cl}$
5. Mass of individual atoms is expressed in *amu* - atomic mass units - 1/12 the mass of single a ^{12}C isotope (1.661×10^{-27} kg)
6. The bottom number in the Periodic Table for each element is the *atomic weight* i.e. the weighted average of the atomic masses of all of the isotopes
 - a. ^{12}C - 98.9%, ^{13}C - 1.1%, ^{14}C - ~0%
 $(0.989)(12) + (0.011)(13) + (0)(14) = 12.011$ amu
 - b. ^{35}Cl - 75.78%, ^{37}Cl - 24.22%
 $(0.7578)(35) + (0.242)(37) = 35.48$ amu
 - c. ^{50}Cr - 4.345%, ^{52}Cr - 83.789%, ^{53}Cr - 9.501%, ^{54}Cr - 2.315%
 $(0.04345)(50) + (0.83789)(52) + (0.09501)(53) + (0.02315)(54) = 52.03$ amu

C. Electrons

1. Electrostatic attraction (i.e. electromagnetic attraction) occurs between the nucleus and orbiting electrons

2. The number of electrons determines an atom's chemical reactivity i.e., the number of electrons and their arrangement around the nucleus determines an atom's chemical reactivity
3. In the elemental state atoms are electrically neutral, i.e., # protons = # electrons
4. Ions - atoms with an electrical charge
 - a. *Ions* - atoms that have either gained or lost electrons
 - b. *Cations* - have lost electrons, since (#electrons < #protons) have net positive charge
 - c. *Anions* - have gained electrons, since (#electrons > #protons) have net negative charge
5. Substance cannot gain protons, they can only gain or lose electrons
6. Examples
 - a. Mg and Mg²⁺
 - b. As and As³⁻
 - c. NH₄⁺ and SO₄²⁻ - molecular ions
7. The Octet Rule and the concept of "isoelectronic"
 - a. Isoelectronic: having the same number of electrons and hence, the same electron configuration, as another element
 - b. The Octet rule: elements "want" to be isoelectronic with the nearest noble gas
 - c. Elements become isoelectronic with the nearest noble gas by gaining or losing electrons (i.e. by forming cations or anions)
 - d. The number of electrons an atom gains or loses depends on how far away it is from the nearest noble gas

D. Atomic symbols - ${}^{\text{am}}_{\text{an}}\text{X}^{\text{charge}}$

IV. The Periodic Table

A. Every element is represented by a symbol

1. Sometimes the name is obvious from the symbol, e.g., O - oxygen, H - hydrogen
2. For some elements the symbol is derived from a non-English language like Latin or German, e.g., K - potassium (kalium), Fe - iron (ferrum), gold - Au (aurum), tungsten - W (wolfram, Swedish)

B. Elements are arranged in rows and columns in sequence of increasing atomic number

1. Rows = periods
2. Columns = groups

C. Groups are "chemical families," exhibit similar chemical behavior

1. Group 1A - alkali metals
2. Group 2A - alkali earth metals
3. Group 6A - chalcogens
4. Group 7A - halogens
5. Group 8A - Noble Gases
6. Also lanthanides and actinides

D. Metals, nonmetals, and metalloids

1. *Metals* - elements characterized by the tendency to give up electrons in reactions (i.e. form cations), by good thermal and electrical conductivity, usually lustrous, malleable, ductile
2. *Nonmetals* - elements characterized by the tendency to gain electrons in reactions (i.e. form anions) and by a lack of other metallic properties
3. *Metalloids (semi-metals)* - elements characterized by the some of the properties of both metals and nonmetals

V. Molecules, molecular compounds, and ionic compounds

A. Molecules - two or more atoms held tightly together by chemical bonds (electrostatic attraction)

B. Chemical formula - molecules are represented using a shorthand notation indicating the number and types of bonds in a compound

1. Empirical formulas: simplest whole number ratio of atoms in a molecule
 - a. Only used in elemental analysis
2. Molecular formulas: provides the number and types of atoms in a compound
3. Shortcomings: 2 or more compounds may have the same molecular formula but with the atoms arranged differently - structural isomers e.g. ethanol vs. dimethyl ether
4. Structural formulas show atoms and connectivities
 - a. Structural formulas

- b. Condensed structural formulas
- c. Backbone formulas
- d. Line formulas

C. Three types of chemical bonds, each resulting in a different type of compound

1. Metallic bonds: between M-M, "electron sea" model; metallic compounds
2. Ionic bonds: between cations and anions, usually between M-NM, bonding electrons not shared equally; ionic compounds
3. Covalent bonds: between NM-NM, bonding electrons shared more or less equally; covalent (molecular) compounds

D. Metallic compounds

1. Have metallic bonds
2. Includes pure metals and alloys
3. "Electron sea" model

E. Ionic compounds

1. Held together by ionic bonds
2. Includes salts and most rocks and minerals

F. Molecular compounds

1. Held together by covalent bonds
2. Includes organic compounds and polymers
3. Polymer - a molecular chain of many smaller repeating units called monomers

VI. Nomenclature of ionic compounds, covalent compounds, and acids and bases

A. Notes

1. Ionic compounds are those that contain ionic bonds; between metals and nonmetals, *or* compounds that contain polyatomic ions
2. We will learn the IUPAC rules of nomenclature (systematic nomenclature) for ionic and covalent compounds
3. Do not worry about the Stock (old) nomenclature system (e.g. ous, ic, ate, ite, hypo, per)
4. The nomenclature of organic compounds has its own set of rules that will only be discussed briefly in this class (you'll have to wait until Chem 2310)

B. Naming ionic compounds

1. Naming monatomic cations

- a. Group 1 and 2 metals have only one cation: element name + "ion"
- b. Transition metals and p-block metals generally have more than one cation: element name + (charge in Roman numerals) + "ion"
 - i. Can use the Periodic Table to help predict the charges of Group 1A, 2A, 5A, 6A and 7A (single atom) ions
 - ii. Substances "want" to be isoelectronic with the nearest Noble Gas
 - iii. Metals lose electrons, nonmetals gain electrons

2. Naming monatomic anions: element name - "end" + "ide"

3. Polyatomic ions - see Table 2.5, p. 64

- a. Polyatomic cations: Hg_2^{2+} mercury (I), NH_4^+ ammonium
- b. Some common polyatomic anions (not an exhaustive list)

SO_4^{2-}	sulfate
HSO_4^-	hydrogen sulfate (bisulfate)
SO_3^{2-}	sulfite
HSO_3^-	hydrogen sulfite (bisulfite)
NO_3^-	nitrate
NO_2^-	nitrite
OH^-	hydroxide
CO_3^{2-}	carbonate
HCO_3^-	hydrogen carbonate (bicarbonate)
CrO_4^{2-}	chromate
MnO_4^-	permanganate
$\text{C}_2\text{H}_3\text{O}_2^-$	acetate
PO_4^{3-}	phosphate
HPO_4^{2-}	hydrogen phosphate
H_2PO_4^-	dihydrogen phosphate
CN^-	cyanide

ClO_4^-	perchlorate
ClO_3^-	chlorate
ClO_2^-	chlorite
ClO^-	hypochlorite

c. Note that many of these are "oxyanions"

4. Ionic compound names

- Cations are named first, anions are named second
- When naming the compound, drop the "ion" portion of the cation name when adding the cation name to the anion name

5. Formulas of ionic compounds

- Molecular formula - a shorthand notation indicating the types and numbers of atoms in a compound
 - By convention the molecular formulae of ionic compounds list cations first, then anions
 - Shortcomings: 2 or more compounds may have the same molecular formula but with the atoms arranged differently - structural (constitutional) isomers e.g. ethanol vs. dimethyl ether
- Compounds are electrically neutral, so there must be a balance between the net positive charge of the cations and the net negative charge of the anions
- Note: you must pay attention to charge when writing the formulas of ionic compounds, worry about mass balance later

6. Going from molecular formulas to names

- Is it ionic?
- Write the symbols for the ions
- Name the ions
- Combine the ion names to name the compound
- Examples: FeCl_3 , NaNO_3 , $(\text{NH}_4)_2\text{SO}_4$, Ba_3P_2

7. Going from names to molecular formulas

- Is it ionic?
- Write the symbols for the ions

- c. Find the LCD of the ion charges
 - d. Write the formula for the neutral compound
 - e. Examples: Potassium chlorate, aluminum sulfite, sodium acetate, rubidium oxide
8. "Cross-multiplying" rule and warning
 9. A word about the Stock system and the intermixture of nonmetal nomenclature prefixes in ionic nomenclature
- C. Binary molecular compounds - contain covalent bonds; between nonmetals and nonmetals

1. Binary compounds consist of two elements
2. Element order established by convention (i.e., history): left to right and bottom to top
 1. B - Si C - Sb As P N - H - Te Se S - I Br Cl - O - F
 2. Only exceptions to rule are H and O
3. Naming rules
 - a. Compound name has elements in same order as molecular formula
 - b. First element: exact name
 - c. Second element: "ide" suffix (named as if it is an anion)
 - d. Prefixes denote numbers of atoms in compound (Table 2.7, p. 72; these result in the subscripts found in molecular formulae):

1	mono
2	di
3	tri
4	tetra
5	penta
6	hexa
7	hepta
8	octa
9	nona
10	deca

- i. Exception 1: first element by itself is never "mono" e.g. nitrogen dioxide

- ii. Chop "o" / "a" from prefix is element name begins with a vowel
e.g. carbon monoxide

4. Examples

- a. XeF_6 , KrF_2 , ICl_5 , N_4S_4 , P_2O_5 , NO , N_2O
- b. Iodine heptafluoride, dinitrogen pentoxide, tetraphosphorus decoxide

D. Acids and bases

- Acids are substances that can donate a hydrogen ion
- Bases are substances that can donate a hydroxide ion
- The names of acids and bases are based on common accepted names and not on the systematic IUPAC nomenclature
- Some typical acids: carbonic acid, nitrous acid, nitric acid, phosphoric acid, sulfurous acid, sulfuric acid, hypochlorous acid, chlorous acid, chloric acid, perchloric acid, hydrofluoric acid, hydrochloric acid, hydrobromic acid, hydroiodic acid
- Some typical acids:

nitric acid	HNO_3
nitrous acid	HNO_2
sulfuric acid	H_2SO_4
sulfurous acid	H_2SO_3
hydrofluoric acid	HF
hydrochloric acid	HCl
hydrobromic acid	HBr
hydroiodic acid	HI
carbonic acid	H_2CO_3
phosphoric acid	H_3PO_4
perchloric acid	HClO_4
chloric acid	HClO_3
chlorous acid	HClO_2
hypochlorous acid	HClO

6. Some typical bases

lithium hydroxide	LiOH
sodium hydroxide	NaOH
potassium hydroxide	KOH
rubidium hydroxide	RbOH

cesium hydroxide	CsOH
calcium hydroxide	Ca(OH) ₂
strontium hydroxide	Sr(OH) ₂
barium hydroxide	Ba(OH) ₂
ammonium hydroxide	NH ₄ OH
ammonia	NH ₃

VII. Organic nomenclature

A. All organic compounds consist of chains of one or more carbon atoms

1. All organic compounds contain both carbon and hydrogen atoms

B. The proper (systematic, or IUPAC) name of organic compounds states two pieces of information about the compound: the class of the compound, as determined by its functional group, and the number of carbon atoms in the longest chain containing the most important functional group

C. The first part of every organic name conveys information about the number of carbon atoms in the longest chain

1. Chain length and names

1	CH ₃ -	methyl
2	C ₂ H ₅ -	ethyl
3	C ₃ H ₇ -	propyl
4	C ₄ H ₉ -	butyl
5	C ₅ H ₁₁ -	pentyl
6	C ₆ H ₁₃ -	hexyl
7	C ₇ H ₁₅ -	heptyl
8	C ₈ H ₁₇ -	octyl
9	C ₉ H ₁₉ -	nonyl
10	C ₁₀ H ₂₁ -	decyl

D. The second part of the name tells the chemical family to which the compound belongs

1. Organic compounds are generally classed by their functional groups, atoms or molecules which determine the general chemistry of the compounds: note that in most cases naming the compound is based in naming the compound as an alkane and then substituting the appropriate suffix for the class of compounds

family	family name	functional groups
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alkane	ane	C-C single bonds
alkene	ene	C-C double bonds
alkyne	yne	C-C triple bonds
aromatic		cyclic compounds with multiple double bonds
alcohol	ol	-OH group
ether	ether	C-O-C
aldehyde	al	-CHO
ketone	one	C-CO-C
carboxylic acid	oic acid	-COOH
ester	alcohol name + acid name changed to "ate"	COOC
amine	amine	organic substituted ammonia
amide	amide	CONH

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