Chapter 2: The Components of Matter

2.1 Elements, Compounds, and Mixtures: An Atomic Overview

2.2 The Observations That Led to an Atomic View of Matter

2.3 Dalton’s Atomic Theory

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2.8 Compounds: Formulas, Names, and Masses

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Definitions for Components of Matter

**Element** - the simplest type of substance with unique physical and chemical properties. An element consists of only one type of atom. It cannot be broken down into any simpler substances by physical or chemical means.

**Molecule** - a structure that consists of two or more atoms that are chemically bound together and thus behaves as an independent unit.

**Compound** - a substance composed of two or more elements which are chemically combined.

**Mixture** - a group of two or more elements and/or compounds that are physically intermingled.

The law of mass conservation.

The total mass of substances does not change during a chemical reaction.

\[ \text{reactant 1} \quad + \quad \text{reactant 2} \quad = \quad \text{product} \]

\[ \text{total mass}_{\text{reactants}} \quad = \quad \text{total mass}_{\text{products}} \]

reactant 1: calcium oxide \[ \text{CaO} \]

reactant 2: carbon dioxide \[ \text{CO}_2 \]

product: calcium carbonate \[ \text{CaCO}_3 \]

56.08 g + 44.01 g = 100.09 g

Law of Mass Conservation

The total mass of substances does not change during a chemical reaction.
Law of Definite (or Constant) Composition

No matter the source, a particular compound is composed of the same elements in the same parts (fractions) by mass.

Calcium carbonate

<table>
<thead>
<tr>
<th>Analysis by Mass (grams/20.0 g)</th>
<th>Mass Fraction (parts/100 part)</th>
<th>Percent by Mass (parts/100 parts)</th>
</tr>
</thead>
<tbody>
<tr>
<td>8.0 g calcium</td>
<td>0.40 calcium</td>
<td>40% calcium</td>
</tr>
<tr>
<td>2.4 g carbon</td>
<td>0.12 carbon</td>
<td>12% carbon</td>
</tr>
<tr>
<td>9.6 g oxygen</td>
<td>0.48 oxygen</td>
<td>48% oxygen</td>
</tr>
<tr>
<td>20.0 g</td>
<td>1.00 part by mass</td>
<td>100% by mass</td>
</tr>
</tbody>
</table>

Figure 2.2

Law of Multiple Proportions

If elements A and B react to form two compounds, the different masses of B that combine with a fixed mass of A can be expressed as a ratio of small whole numbers.

Example: Carbon Oxides I & II

Carbon Oxide I: 57.1% oxygen and 42.9% carbon
Carbon Oxide II: 72.7% oxygen and 27.3% carbon

Assume that you have 100 g of each compound.

In 100 g of each compound: g O = 57.1 g for oxide I & 72.7 g for oxide II

\[ \frac{g \text{ O}}{g \text{ C}} = \frac{57.1}{42.9} = 1.33 \]

Carbon oxide II

\[ \frac{g \text{ O}}{g \text{ C}} = \frac{72.7}{27.3} = 2.66 \]

Carbon oxide I

\[ \frac{2.66 \text{ g O in B}}{1.33 \text{ g O in A}} = \frac{2}{1} \]

Dalton’s Atomic Theory

The Postulates

1. All matter consists of atoms.
2. Atoms of one element cannot be converted into atoms of another element.
3. Atoms of an element are identical (in mass** and other properties) and are different from atoms of any other element.
4. Compounds result from the chemical combination of a specific ratio of atoms of different elements.

Mass conservation

Atoms cannot be created or destroyed postulate 1

or converted into other types of atoms. postulate 2

Since every atom has a fixed “mass”, postulate 3
during a chemical reaction atoms are simply recombined differently; therefore there is no mass change overall.

Definite composition

Atoms are combined in compounds in specific ratios postulate 3

and each atom has a specific “mass”. postulate 4

So each element has a fixed fraction of the total mass in a compound.

Multiple proportions

Atoms of an element have the same “mass” postulate 3

and atoms are indivisible. postulate 4

So when different numbers of atoms of elements combine, they must do so in ratios of small, whole numbers.
Experiments to determine the properties of cathode rays.

Millikan's oil-drop experiment for measuring an electron's charge.

Millikan used his findings to also calculate the mass of an electron.

Rutherford's α-scattering experiment and discovery of the atomic nucleus.

General features of the atom today.

Properties of the Three Key Subatomic Particles

<table>
<thead>
<tr>
<th>Name (Symbol)</th>
<th>Charge (Relative)</th>
<th>Absolute (C)*</th>
<th>Relative (amu)†</th>
<th>Absolute (g)</th>
<th>Location in the Atom</th>
</tr>
</thead>
<tbody>
<tr>
<td>Proton (p⁺)</td>
<td>1+</td>
<td>+1.60218 x 10⁻¹⁹</td>
<td>1.60727</td>
<td>1.67262 x 10⁻²⁴</td>
<td>Nucleus</td>
</tr>
<tr>
<td>Neutron (n⁻)</td>
<td>0</td>
<td>0</td>
<td>1.00866</td>
<td>1.67493 x 10⁻²⁴</td>
<td>Nucleus</td>
</tr>
<tr>
<td>Electron (e⁻)</td>
<td>-1</td>
<td>-1.60218 x 10⁻¹⁹</td>
<td>9.10939 x 10⁻²⁸</td>
<td>9.10939 x 10⁻²⁸</td>
<td>Outside Nucleus</td>
</tr>
</tbody>
</table>

* The coulomb (C) is the SI unit of charge.
† The atomic mass unit (amu) equals 1.66054 x 10⁻²⁴ g.
Atomic Symbols, Isotopes, Numbers

\[ X \] – The Symbol of the Atom or Isotope

- \( X \): atomic symbol of the element
- \( A \): nuclide (or mass) number; \( A = Z + N \)
- \( Z \): atomic number (the number of protons in the nucleus)
- \( N \): number of neutrons in the nucleus

Isotopes = atoms of an element with the same number of protons, but a different number of neutrons

Average Atomic Mass

Since elements are found in nature as mixtures of isotopes, and each isotope is found in a fixed amount in nature, and rarely are these amounts equal among the given isotopes of an element, we must have a way to take this into account when talking about a naturally occurring element. Enter Average Mass:

\[
\text{Ave Atomic Mass} = \sum_{\text{all isotopes}} (f_a \text{ Isotope}_i \times \text{mass Isotope}_i)
\]

\[ f_a = \text{fractional abundance = } \frac{\% \text{ Abundance isotope}}{100} \]

The Modern Reassessment of the Atomic Theory

1. All matter is composed of atoms. The atom is the smallest body that retains the unique identity of the element.
2. Atoms of one element cannot be converted into atoms of another element in a chemical reaction. Elements can only be converted into other elements in nuclear reactions.
3. All atoms of an element have the same number of protons and electrons, which determines the chemical behavior of the element. Isotopes of an element differ in the number of neutrons, and thus in mass number. A sample of the element is treated as though its atoms have an average mass.
4. Compounds are formed by the chemical combination of two or more elements in specific ratios.

Some metals, metalloids, and nonmetals
Transferring electrons from the atoms of one element to those of another results in an ionic compound.

Factors that influence the strength of ionic bonding.

The relationship between ions formed and the nearest noble gas.

Formation of a covalent bond between two H atoms

Elements that occur as molecules

Molecules and Ions

- **Molecular** – the basic unit of an element or covalent compound, consisting of two or more atoms bonded by the sharing of electrons.

- Most covalent substances consist of molecules.

- **Ions** – a single atom or covalently bonded group of atoms that has an overall electrical charge.

- There are no molecules in an ionic compound.
A polyatomic ion consists of two or more atoms covalently bonded together and has an overall charge. In many reactions the polyatomic ion will remain together as a unit.

Nomenclature of Inorganic Compounds

Naming Simple Inorganic Compounds
Dr. Carl Hoeger
profcah

Nomenclature-Basics

Two general classes of inorganic molecules:
1. Those containing a metal cation and a nonmetal anion OR a metal cation and a polyatomic anion (ionic compounds);
2. Those containing only nonmetals (molecular compounds; Type III compounds)

Two different types of the first class:
a) Those whose metal has only one possible positive oxidation state (Type I compounds);
b) Those whose metal has more than one possible positive oxidation state (Type II compounds)
The nomenclature scheme reflects the compounds class/type

Types of Chemical Formulas

A chemical formula is comprised of element symbols and numerical subscripts that show the type and number of each atom present in the smallest unit of the substance.

An empirical formula indicates the relative number of atoms of each element in the compound. It is the simplest type of formula.
The empirical formula for hydrogen peroxide is HO.

A molecular formula shows the actual number of atoms of each element in a molecule of the compound.
The molecular formula for hydrogen peroxide is H₂O₂.

A structural formula shows the number of atoms and the bonds between them, that is, the relative placement and connections of atoms in the molecule.
The structural formula for hydrogen peroxide is H-O-O-H.

Nomenclature-History

Early: Compounds named by those who discovered them; usually had some historic significance
• Oil of Vitrol
• Blue Vitrol
• Laughing gas

Problem: No real system; region/researcher specific; needed systematic, well-defined nomenclature system

Answer: IUPAC development of standardized nomenclature schemes; some trivial names accepted and retained

“Type I” Elements

The carbonate ion in calcium carbonate.
**Chap 2: Matter and Nomenclature**

### “Type II” Elements

**Nomenclature Periodic Table**

- **Nomenclature: Cation Basics**
  - **Type I**: Use the name of the element; if you are talking about an ion by itself put ‘ion’ after the name:
    - Na$^+$ = sodium ion; Mg$^{2+}$ = magnesium ion
  - Special cations: NH$_4^+$ = ammonium ion; H$_3$O$^+$ = hydronium ion

- **Type II**: Use the name of the element AND its oxidation state (or charge); use roman numerals in parentheses to denote this; if you are talking about an ion by itself put ‘ion’ after the name:
  - Cu$^{2+}$ = copper (II) ion; Cr$^{6+}$ = chromium (VI) ion

- **Type III**: Use the name; denote number of that element present using greek prefixes (more later)

### “Type III” Elements: Nonmetals

**Nomenclature: Anion Basics**

- **For ALL**: Take the root name of the element and add -ide to the end; if you are talking about an ion by itself put ‘ion’ after the name:
  - Cl$^-$ = chloride ion; S$^{2-}$ = sulfide ion; P$^{3-}$ = phosphate ion

- **Type III Compounds ONLY**: Add to the name of the second element a prefix denoting how many of that element there are:
  - XO$_2$ would be “X dioxide”

- **POLYATOMIC ANIONS**: These have their own nomenclature you need to know

### Nomenclature: Polyatomic Anions

- Learn these polyatomic anions NOW:
  - HCO$_3^-$ = Hydrogen carbonate
  - HS$^-$/H$_2$S$^-$ = Hydrogen sulfide
  - HSO$_3^-$ = Hydrogen sulfate
  - H$_2$PO$_4^-$ = Dihydrogen phosphate
  - CO$_2$ = Carbonate
  - SO$_4^{2-}$ = Sulfate
  - CO$_3^{2-}$ = Carbonate
  - SO$_3^{2-}$ = Sulfite
  - C$_2$O$_4^{2-}$ = Oxalate
  - NO$_3^-$ = Nitrate
  - NO$_2^-$ = Nitrite
  - OH$^-$ = Hydroxide
  - MnO$_4^-$ = Permanganate
  - CO$_3^{2-}$ = Carbonate
  - SO$_4^{2-}$ = Sulfate
  - SO$_3^{2-}$ = Sulfite
  - C$_2$O$_4^{2-}$ = Oxalate
  - Br$^-$/Br$_2$ = Bromide
  - BrO$^-$/BrO$_2^-$ = Hypobromite
  - BrO$_3^-$ = Bromate
  - BrO$_4^-$ = Perbromate
  - SCN$^-$ = Thiocyanate
  - CN$^-$ = Cyanide
  - PO$_4^{3-}$ = Phosphate
  - PO$_3^{3-}$ = Phosphite
  - C$_2$H$_3$O$_2^-$ = Acetate

9/4/11
**Nomenclature: Polyatomic Anions**

**Learn these "-ate" polyatomic anions NOW:**

<table>
<thead>
<tr>
<th>Prefix</th>
<th>Root</th>
<th>Suffix</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>per</td>
<td>root</td>
<td>ate</td>
<td>ClO₄⁻</td>
</tr>
<tr>
<td>root</td>
<td>ate</td>
<td>ClO₄⁻</td>
<td>perchlorate</td>
</tr>
<tr>
<td>root</td>
<td>ite</td>
<td>ClO₃⁻</td>
<td>chloride</td>
</tr>
<tr>
<td>hypo</td>
<td>root</td>
<td>ite</td>
<td>ClO₂⁻</td>
</tr>
<tr>
<td>hypo</td>
<td>root</td>
<td>ate</td>
<td>ClO₁⁻</td>
</tr>
</tbody>
</table>

**Summary: Naming oxoanions**

- **Prefix**
- **Root**
- **Suffix**
- **Example**

**Nomenclature: Oxyanions**

Systematic nomenclature based on oxoacids of origin;
- End in -ate or -ite; may have per- or hypo- as prefixes;
- Learn name, formula, and charge of all -ate ions, then:
  a) If an ion has **ONE LESS** oxygen than the -ate ion, change -ate to -ite (charge does not change);
  b) If an ion has **TWO LESS** oxygens than the -ate ion, change -ate to -ite and add hypo- as a prefix (charge does not change);
  c) If an ion has **ONE MORE** oxygen than the -ate ion, add per- as a prefix (charge does not change).

**Nomenclature: Type I and II Compounds**

Simple: take name of cation and anion and combine:
- **CaCl₂ = calcium chloride**
- **ZnS = zinc sulfide**
- **Ba₃(PO₄)₂ = barium phosphate**
- **CrO₃ = chromium (VI) oxide**
- **PbI₂ = lead (IV) iodide**
- **Fe₂O₃ = iron (III) oxide**

**Nomenclature: Type III Compounds**

1. **START** by naming as if it were a Type I compound:
   - **N₂F₄: nitrogen fluoride**
2. **THEN** add a greek prefix to tell how MANY of each element there are:
   - **N₅F₁₂: dinitrogen tetrafluoride**
   - **Note:** if there is only one of the first element, the prefix mono is **NOT** used:
   - **SO₃: sulfur dioxide NOT monosulfur dioxide**

**Determining cation charge for Type II ions**

- Based on **knowing your anions and their charges**
- A compound is neutral and has a charge of zero
- The sum of all the oxidation states in a compound must equal the overall charge on the ion or molecule.
- Set up and solve for the unknown algebraically:

\[
\text{charge of } \text{Fe}_2\text{O}_3 = 2 \times \text{charge of iron} + (3 \times \text{charge of oxygen}) = \text{charge of } \text{Fe}_2\text{O}_3
\]

1. **START** by naming as if it were a Type I compound:
   - **N₂F₄: nitrogen fluoride**
2. **THEN** add a greek prefix to tell how MANY of each element there are:
   - **N₅F₁₂: dinitrogen tetrafluoride**
   - **Note:** if there is only one of the first element, the prefix mono is **NOT** used:
   - **SO₃: sulfur dioxide NOT monosulfur dioxide**
Try the following:

- \( \text{C}_2\text{O}_4 \)
- \( \text{Na}_2\text{SO}_3 \)
- \( \text{MgO} \)
- \( \text{CrO}_2 \)
- \( \text{P}_5\text{O}_{10} \)
- \( \text{AsO}_3 \)

Give yourself five minutes.

Answers:

- \( \text{C}_2\text{O}_4 \) dicarbon tetroxide
- \( \text{Na}_2\text{SO}_3 \) sodium sulfite
- \( \text{MgO} \) magnesium oxide
- \( \text{CrO}_2 \) chromium (IV) sulfide
- \( \text{P}_5\text{O}_{10} \) pentaphosphorous decoxide
- \( \text{AsO}_3 \) arsenic trioxide

Answers:

- \( \text{SeO}_4 \) selenium tetroxide
- \( \text{Al}_2\text{Cr}_2\text{O}_7 \) aluminum dichromate
- \( \text{Mo}\text{CO}_3\text{Cl}_2 \) molybdenum (IV) carbonate
- \( \text{Rh}\text{NO}_3\text{Cl}_3 \) rhodium (V) nitrate
- \( \text{H}_2\text{O} \) dihydrogen monoxide
- \( \text{N}_2\text{O} \) dinitrogen monoxide

1. Some compounds have been known and used for so long that their trivial (or common names) have become accepted by the IUPAC as official:

- \( \text{H}_2\text{O} = \text{water} \)
- \( \text{CH}_4 = \text{methane} \)
- \( \text{NH}_3 = \text{ammonia} \)
- \( \text{SiH}_4 = \text{silane} \)
- \( \text{N}_2\text{H}_4 = \text{hydrazine} \)

2. Oxides of some nonmetals will sometimes be named as if they were Type II compounds:

- \( \text{P}_2\text{O}_5 = \text{phosphorous (V) oxide} \)
- \( \text{SeO}_3 = \text{selenium (VI) oxide} \)

Note: this is an older system that is slowly being phased out and we will not use.

3. For Type II metals with only two common oxidation states an older, Latin system was once used; while it is not employed very often it is useful to know some simple rules regarding it. It is sometimes called the “-ous/-ic” system, where the LOWER charged cation will be denoted by the latin root name for that element plus -ous and the HIGHER charged cation will be denoted by the latin root name for that element plus -ic:

- \( \text{Fe}^{2+} = \text{ferrous ion} \)
- \( \text{Fe}^{3+} = \text{ferric ion} \)
- \( \text{Sn}^{2+} = \text{stannous ion} \)
- \( \text{Sn}^{4+} = \text{stannic ion} \)

Not used much but still can be found being employed; important in lab!

4. Compounds containing hydrogen listed as the first element are acids and named differently, as are many carbon compounds.

<table>
<thead>
<tr>
<th>Element</th>
<th>Ion Formula</th>
<th>Systematic Name</th>
<th>Common Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>Chromium</td>
<td>( \text{Cr}^{2+} )</td>
<td>chromium(II)</td>
<td>chromous</td>
</tr>
<tr>
<td></td>
<td>( \text{Cr}^{3+} )</td>
<td>chromium(III)</td>
<td>chromic</td>
</tr>
<tr>
<td>Cobalt</td>
<td>( \text{Co}^{2+} )</td>
<td>cobalt(II)</td>
<td>cobalt(III)</td>
</tr>
<tr>
<td>Copper</td>
<td>( \text{Cu}^{+} )</td>
<td>copper(I)</td>
<td>cuprous</td>
</tr>
<tr>
<td></td>
<td>( \text{Cu}^{2+} )</td>
<td>copper(II)</td>
<td>cupric</td>
</tr>
<tr>
<td>Iron</td>
<td>( \text{Fe}^{2+} )</td>
<td>iron(II)</td>
<td>ferrous</td>
</tr>
<tr>
<td></td>
<td>( \text{Fe}^{3+} )</td>
<td>iron(III)</td>
<td>ferric</td>
</tr>
<tr>
<td>Lead</td>
<td>( \text{Pb}^{2+} )</td>
<td>lead(II)</td>
<td>plumbous</td>
</tr>
<tr>
<td></td>
<td>( \text{Pb}^{4+} )</td>
<td>lead(IV)</td>
<td>plumbic</td>
</tr>
<tr>
<td>Mercury</td>
<td>( \text{Hg}^{2+} )</td>
<td>mercury(II)</td>
<td>mercurous</td>
</tr>
<tr>
<td></td>
<td>( \text{Hg}^{4+} )</td>
<td>mercury(I)</td>
<td>mercuric</td>
</tr>
<tr>
<td>Tin</td>
<td>( \text{Sn}^{4+} )</td>
<td>tin(IV)</td>
<td>stannous</td>
</tr>
<tr>
<td></td>
<td>( \text{Sn}^{2+} )</td>
<td>tin(II)</td>
<td>stannic</td>
</tr>
</tbody>
</table>

*Listed alphabetically by metal name; the ions in boldface are most common.
Acid Nomenclature
An acid is a compound that contains an acidic hydrogen (a H that ionizes in water):

\[ HX(aq) \rightarrow H^+(aq) + X^-(aq) \]

Not all hydrogens are ionizable; by convention, to indicate when one IS, we begin the chemical formula with it:

\[ HC_2H_3O_2 \] (acetic acid) has four hydrogens but only ONE is acidic (underlined for emphasis)

There are two types of acids: binary and oxoacids

Binary Acid Nomenclature
For binary acids (HX(aq), where X does NOT contain oxygen; it MUST be in aqueous solution):
Add hydro...ic acid to the root of the element X:

\[ H_2S(aq) = \text{hydrosulfuric acid}; HI(aq) = \text{hydroiodic acid}; HN_3(aq) = \text{hydrazoic acid} \]

Note: if phase is NOT aqueous, name as if it were a Type I compound (even though they are considered to be molecular in nature):

\[ H_2S(g) = \text{hydrogen sulfide}; HI(g) = \text{hydrogen iodide}; HN_3(l) = \text{hydrogen azide} \]

Oxoacid Nomenclature
If the acid contains oxygen (HOX or HXO) its name will be a function of the name of the anion portion (i.e. the portion remaining after all the acidic H’s have been removed: OX or XO_n):

- Determine the anion name:
  - If the name ends in -ate, replace -ate with -ic acid
  - If the name ends in -ite, replace -ite with -ous acid

Leave prefixes alone; the root name is sometimes tweaked to aid in pronunciation

\[ H_2SO_4(aq) \rightarrow \text{anion is SO}_4^{2-} \rightarrow \text{acid name is sulfuric acid} \]
\[ HClO(aq) \rightarrow \text{anion is ClO}^- \rightarrow \text{acid name is hypochlorous acid} \]

Acid Nomenclature: Practice

Try the following:

\[ H_2C_2O_4 \quad H_2SO_3 \quad HCN \quad H_2CrO_4 \quad H_3PO_4 \quad H_2Se \quad HIO_4 \quad H_2CO_3 \quad HCN \quad H_2CrO_4 \quad H_3PO_4 \quad H_2Se \]

Give yourself five minutes

Acid Nomenclature: Practice

Answers:

\[ H_2C_2O_4 \quad \text{oxalic acid} \]
\[ H_2SO_3 \quad \text{sulfurous acid} \]
\[ HCN \quad \text{hydrocyanic acid} \]
\[ H_2CrO_4 \quad \text{chromic acid} \]
\[ H_3PO_4 \quad \text{phosphoric acid} \]
\[ H_2Se \quad \text{hydroselenic acid} \]
\[ HIO_4 \quad \text{periodic acid} \]
\[ H_2CO_3 \quad \text{carbonic acid} \]
\[ HF \quad \text{hydrofluoric acid} \]
\[ HClO_3 \quad \text{acetic acid} \]
\[ HNO_3 \quad \text{nitric acid} \]
\[ HNO_2 \quad \text{nitrous acid} \]
Mixtures

Heterogeneous mixtures: has one or more visible boundaries between the components (sand in water; oil in water).

Homogeneous mixtures: has no visible boundaries because the components are mixed as individual atoms, ions, and molecules (sugar in water; alcohol in water).

Solutions: A homogeneous mixture is also called a solution. Solutions in water are called aqueous solutions, and are very important in chemistry. Although we normally think of solutions as liquids, they can exist in all three physical states.

Basic Separation Techniques

Filtration: Separates components of a mixture based upon differences in particle size. Filtration usually involves separating a precipitate from solution.

Crystallization: Separation is based upon differences in solubility of components in a mixture.

Distillation: separation is based upon differences in volatility.

Extraction: Separation is based upon differences in solubility in different solvents (major material).

Chromatography: Separation is based upon differences in solubility in a solvent versus a stationary phase.

End Chapter 2