Chapter 2

The Components of Matter

Chapter 2: The Components of Matter

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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.

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**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.

**Law of Mass Conservation**

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

\[
\text{reactant 1} + \text{reactant 2} \rightarrow \text{product}
\]

\[
\text{total mass} = \text{total mass}
\]

caesium oxide + carbon dioxide → calcium carbonate

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<tr>
<th>CaO</th>
<th>+</th>
<th>CO₂</th>
<th>→</th>
<th>CaCO₃</th>
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</thead>
<tbody>
<tr>
<td>56.08 g</td>
<td>+</td>
<td>44.01 g</td>
<td>→</td>
<td>100.09 g</td>
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</table>
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>
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<tr>
<th>Analysis by Mass (grams/20.0 g)</th>
<th>Mass Fraction (parts/1.00 part)</th>
<th>Percent by Mass (parts/100 parts)</th>
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<tbody>
<tr>
<td>8.0 g calcium</td>
<td>0.40 calcium</td>
<td>40% calcium</td>
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<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>
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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
- g C = 42.9 g for oxide I & 27.3 g for oxide II

\[
\frac{g \text{ O}}{g \text{ C}} = \frac{57.1}{42.9} = 1.33 \\
\frac{g \text{ O}}{g \text{ C}} = \frac{72.7}{27.3} = 2.66
\]

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

Chap 2: Matter and Nomenclature

Mass conservation

Atoms cannot be created or destroyed or converted into other types of atoms. Since every atom has a fixed “mass”, during a chemical reaction atoms are simply recombined differently; therefore there is no mass change overall.
**Dalton’s Atomic Theory**

explains the mass laws

**Definite composition**

Atoms are combined in compounds in specific postulate 3 ratios

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

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

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**Dalton’s Atomic Theory**

explains the mass laws

**Multiple proportions**

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

and atoms are indivisible. postulate 1

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.

<table>
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<th>OBSERVATION</th>
<th>CONCLUSION</th>
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<tr>
<td>1. Ray bends in magnetic field</td>
<td>Consists of charged particles</td>
</tr>
<tr>
<td>2. Ray bends toward positive plate in electric field</td>
<td>Consists of negative particles</td>
</tr>
<tr>
<td>3. Ray is identical for any cathode</td>
<td>Particles found in all matter</td>
</tr>
</tbody>
</table>

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

Oil droplets fall through hole in positively charged plate
Fine mist of oil sprayed into apparatus
X-rays knock electrons from surrounding air, which stick to droplets
Electrically charged plates influence droplets motion
Observe times droplet’s motion and controls electric field

(1909)
Millikan used his findings to also calculate the mass of an electron.

\[
\text{mass of electron} = \frac{\text{mass}}{\text{charge}} \times \text{charge} = (-5.686 \times 10^{-12} \text{ kg}) \times (-1.602 \times 10^{-19} \text{ C})
\]

\[
= 9.109 \times 10^{-31} \text{ kg} = 9.109 \times 10^{-28} \text{ g}
\]

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

1 in 20,000
General features of the atom today.

- The atom is an electrically neutral, spherical entity composed of a positively charged central nucleus surrounded by one or more negatively charged electrons.
- The atomic nucleus consists of protons and neutrons.

Table 2.2  Properties of the Three Key Subatomic Particles

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<tr>
<th>Name (Symbol)</th>
<th>Charge</th>
<th>Mass</th>
<th>Location in the Atom</th>
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<tr>
<td>Proton (p⁺)</td>
<td>1⁺</td>
<td>1.00727</td>
<td>1.67262 x 10⁻²⁴</td>
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<tr>
<td>Neutron (n⁰)</td>
<td>0</td>
<td>0</td>
<td>1.00866</td>
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<tr>
<td>Electron (e⁻)</td>
<td>1⁻</td>
<td>-1.60218 x 10⁻¹⁹</td>
<td>0.00054858</td>
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</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

\[ ^{A}_{Z}X \] — The Symbol of the Atom or Isotope

\( X \) = atomic symbol of the element

\( A \) = nuclide (or mass) number; \( A = Z + N \)

\( Z \) = atomic number

(\text{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

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

Isotopes have the same atomic number, but a different mass number.
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:

\[
\text{Ave Atomic Mass} = \sum_{\text{for all isotopes}} \left( fa_{\text{Isotope}_i} \right) \left( \text{mass Isotope}_i \right)
\]

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

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

Figure 2.10 The modern periodic table.

Some metals, metalloids, and nonmetals:

- Chromium (Z = 24)
- Copper (Z = 29)
- Cadmium (Z = 48)
- Bismuth (Z = 83)
- Arsenic (Z = 33)
- Antimony (Z = 55)
- Tellurium (Z = 52)
- Sulfur (Z = 16)
- Chlorine (Z = 17)
- Bromine (Z = 35)
- Iodine (Z = 53)
- Carbon (graphite)
- Boron (Z = 5)

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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 formation of an ionic compound.

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.

\[ E_{\text{Coulombic}} \propto \frac{Q_1 \times Q_2}{r_{12}} \]

Attraction increases
The relationship between ions formed and the nearest noble gas.

Formation of a covalent bond between two H atoms

Covalent bonds form when elements share electrons, which usually occurs between nonmetals.
**Molecules and Ions**

**Molecule** – 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.

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

There are no molecules in an ionic compound.

**Elements that occur as molecules**

Have No Fear Of Ice Cold Beer
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.

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 $H_2O_2$.*

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 of Inorganic Compounds

Naming Simple Inorganic Compounds

Dr. Carl Hoeger

profcah

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
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

“Type I” Elements
“Type II” Elements

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“Type III” Elements: Nonmetals

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</tbody>
</table>
Nomenclature: Cation Basics

To name a cation (or the first element in a Type III compound):

✔ **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; Mg²⁺ = magnesium ion
  
  *Special cations*: NH₄⁺ = ammonium ion;
  
  H₃O⁺ = 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:

  Cu²⁺ = copper (II) ion; Cr⁶⁺ = chromium (VI) ion

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

To name a monatomic anion (OR the second element in a binary Type III compound):

- **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\(^{3-}\) = phosphide 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

<table>
<thead>
<tr>
<th>Anion</th>
<th>Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>HCO(_3^-)</td>
<td>Hydrogen carbonate</td>
</tr>
<tr>
<td>HS(^-)</td>
<td>Hydrogen sulfide</td>
</tr>
<tr>
<td>HSO(_4^-)</td>
<td>Hydrogen sulfate</td>
</tr>
<tr>
<td>HSO(_3^-)</td>
<td>Hydrogen sulfite</td>
</tr>
<tr>
<td>H(_2)PO(_4^-)</td>
<td>Dihydrogen phosphate</td>
</tr>
<tr>
<td>ClO(_3^-)</td>
<td>Perchlorate</td>
</tr>
<tr>
<td>ClO(_2^-)</td>
<td>Chlorate</td>
</tr>
<tr>
<td>ClO(^-)</td>
<td>Chlorite</td>
</tr>
<tr>
<td>ClO(_2^-)</td>
<td>Hypochlorite</td>
</tr>
<tr>
<td>BrO(_3^-)</td>
<td>Bromate</td>
</tr>
<tr>
<td>BrO(_2^-)</td>
<td>Bromite</td>
</tr>
<tr>
<td>BrO(^-)</td>
<td>Hypobromite</td>
</tr>
<tr>
<td>IO(_4^-)</td>
<td>Periodate</td>
</tr>
<tr>
<td>IO(_3^-)</td>
<td>Iodate</td>
</tr>
<tr>
<td>C(_2)H(_3)O(_2^-)</td>
<td>Acetate</td>
</tr>
<tr>
<td>NO(_3^-)</td>
<td>Nitrate</td>
</tr>
<tr>
<td>NO(_2^-)</td>
<td>Nitrite</td>
</tr>
<tr>
<td>OH(^-)</td>
<td>Hydroxide</td>
</tr>
<tr>
<td>MnO(_4^-)</td>
<td>Permanganate</td>
</tr>
<tr>
<td>CO(_3^-)</td>
<td>Carbonate</td>
</tr>
<tr>
<td>SO(_3^-)</td>
<td>Sulfate</td>
</tr>
<tr>
<td>SO(_2^-)</td>
<td>Sulfite</td>
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<tr>
<td>C(_2)O(_4^-)</td>
<td>Oxalate</td>
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<tr>
<td>CrO(_3^-)</td>
<td>Chromate</td>
</tr>
<tr>
<td>CrO(_2^-)</td>
<td>Dichromate</td>
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<tr>
<td>SCN(^-)</td>
<td>Thiocyanate</td>
</tr>
<tr>
<td>CN(^-)</td>
<td>Cyanide</td>
</tr>
<tr>
<td>HPO(_4^-)</td>
<td>Hydrogen Phosphate</td>
</tr>
<tr>
<td>PO(_3^-)</td>
<td>Phosphate</td>
</tr>
</tbody>
</table>
Nomenclature: Polyatomic Anions

Learn these “-ate” polyatomic anions NOW:

<table>
<thead>
<tr>
<th>Anion</th>
<th>Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>HCO(^3)⁻</td>
<td>Hydrogen carbonate</td>
</tr>
<tr>
<td>HS⁻</td>
<td>Hydrogen sulfide</td>
</tr>
<tr>
<td>HSO(^3)⁻</td>
<td>Hydrogen sulfite</td>
</tr>
<tr>
<td>HSO(^4)⁻</td>
<td>Dihydrogen phosphate</td>
</tr>
<tr>
<td>ClO(^4)⁻</td>
<td>Perchlorate</td>
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<tr>
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<td>Chlorate</td>
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<td>ClO(^2)⁻</td>
<td>Chlorite</td>
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<tr>
<td>ClO⁻</td>
<td>Hypochlorite</td>
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<tr>
<td>BrO(^3)⁻</td>
<td>Bromate</td>
</tr>
<tr>
<td>BrO(^2)⁻</td>
<td>Bromite</td>
</tr>
<tr>
<td>BrO⁷⁻</td>
<td>Hypobromite</td>
</tr>
<tr>
<td>IO(^4)⁻</td>
<td>Periodate</td>
</tr>
<tr>
<td>IO(^3)⁻</td>
<td>Iodate</td>
</tr>
<tr>
<td>C(_2)H(_3)O(^2)⁻</td>
<td>Acetate</td>
</tr>
<tr>
<td>NO(^3)⁻</td>
<td>Nitrate</td>
</tr>
<tr>
<td>NO(^2)⁻</td>
<td>Nitrite</td>
</tr>
<tr>
<td>OH⁻</td>
<td>Hydroxide</td>
</tr>
<tr>
<td>MnO(^4)⁻</td>
<td>Permanganate</td>
</tr>
<tr>
<td>CO(^3)²⁻</td>
<td>Carbonate</td>
</tr>
<tr>
<td>SO(^2)²⁻</td>
<td>Sulfite</td>
</tr>
<tr>
<td>SO(^3)²⁻</td>
<td>Sulfite</td>
</tr>
<tr>
<td>C(_2)O(^2)²⁻</td>
<td>Oxalate</td>
</tr>
<tr>
<td>CrO(^4)²⁻</td>
<td>Chromate</td>
</tr>
<tr>
<td>Cr(_2)O(^7)²⁻</td>
<td>Dichromate</td>
</tr>
<tr>
<td>SCN⁻</td>
<td>Thiocyanate</td>
</tr>
<tr>
<td>CN⁻</td>
<td>Cyanide</td>
</tr>
<tr>
<td>HPO(^4)²⁻</td>
<td>Hydrogen Phosphate</td>
</tr>
<tr>
<td>PO(^4)³⁻</td>
<td>Phosphate</td>
</tr>
</tbody>
</table>

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!).
Summary: Naming oxoanions

<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>ClO$_4^-$ perchlorate</td>
</tr>
<tr>
<td>root</td>
<td>root</td>
<td>ate</td>
<td>ClO$_3^-$ chlorate</td>
</tr>
<tr>
<td>root</td>
<td>root</td>
<td>ite</td>
<td>ClO$_2^-$ chlorite</td>
</tr>
<tr>
<td>hypo</td>
<td>root</td>
<td>ite</td>
<td>ClO$^-$ hypochlorite</td>
</tr>
</tbody>
</table>

Nomenclature: Type I and II Compounds

Simple: take name of cation and anion and combine:

- $\text{CaCl}_2 = \text{calcium chloride}$
- $\text{ZnS} = \text{zinc sulfide}$
- $\text{Ba}_3(\text{PO}_4)_2 = \text{barium phosphate}$
- $\text{CrO}_3 = \text{chromium (VI) oxide}$
- $\text{PbI}_4 = \text{lead (IV) iodide}$
- $\text{Fe}_2\text{O}_3 = \text{iron (III) oxide}$
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; so for Fe$_2$O$_3$:

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

\[
[2x]+[3(-2)] = 0; \text{ solve to get } x = +3
\]

Nomenclature: Type III Compounds

1. *START* by naming as if it were a Type I compound:

\[
\text{N}_2\text{F}_4: \text{nitrogen fluoride}
\]

2. *THEN* add a greek prefix to tell how MANY of each element there are:

\[
\text{N}_2\text{F}_4: \text{dinitrogen tetrafluoride}
\]

* Note: if there is only one of the first element, the prefix mono is NOT used:

\[
\text{SO}_2: \text{sulfur dioxide NOT monosulfur dioxide}
\]

<table>
<thead>
<tr>
<th>Prefix</th>
<th>Number</th>
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<tbody>
<tr>
<td>mono</td>
<td>1</td>
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<tr>
<td>di</td>
<td>2</td>
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<tr>
<td>tri</td>
<td>3</td>
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<tr>
<td>tetra</td>
<td>4</td>
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<tr>
<td>penta</td>
<td>5</td>
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<tr>
<td>hexa</td>
<td>6</td>
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<td>hepta</td>
<td>7</td>
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<tr>
<td>octa</td>
<td>8</td>
</tr>
<tr>
<td>nona</td>
<td>9</td>
</tr>
<tr>
<td>deca</td>
<td>10</td>
</tr>
</tbody>
</table>
Nomenclature: Practice

Try the following:

- $C_2O_4$  
  $SeO_4$
- $Na_2SO_3$  
  $Al_2Cr_2O_7$  
  (*Give yourself five minutes*)
- $MgO$  
  $Mo(CO_3)_2$
- $CrO_2$  
  $Rh(NO_3)_5$
- $P_5O_{10}$  
  $H_2O$
- $AsO_3$  
  $N_2O$

Nomenclature: Practice

Answers:

- $C_2O_4$  *dicarbon tetroxide*
- $Na_2SO_3$  *sodium sulfite*
- $MgO$  *magnesium oxide*
- $CrS_2$  *chromium (IV) sulfide*
- $P_5O_{10}$  *pentaphosphorous decoxide*
- $AsO_3$  *arsenic trioxide*
Nomenclature: Practice

Answers:

\[ \text{SeO}_4 \]  selenium tetroxide
\[ \text{Al}_2\text{Cr}_2\text{O}_7 \]  aluminum dichromate
\[ \text{Mo}(\text{CO}_3)_2 \]  molybdenum (IV) carbonate
\[ \text{Rh(NO}_3)_5 \]  rhodium (V) nitrate
\[ \text{H}_2\text{O} \]  dihydrogen monoxide
\[ \text{N}_2\text{O} \]  dinitrogen monoxide

Nomenclature: Exceptions (1)

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:

\[ \text{H}_2\text{O} = \text{water} \]
\[ \text{NH}_3 = \text{ammonia} \]
\[ \text{SiH}_4 = \text{silane} \]
\[ \text{N}_2\text{H}_4 = \text{hydrazine} \]

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:

\[
\begin{align*}
\text{Fe}^{2+} & = \text{ferrous ion/Fe}^{3+} = \text{ferric ion; Sn}^{2+} = \text{stannous ion/Sn}^{4+} = \text{stannic ion} \\
\text{Not used much but still can be found being employed; important in lab!}
\end{align*}
\]

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

### Table 2.4: Examples of Metals That Form More Than One Monatomic Ion*

<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>Cr(^{2+})</td>
<td>chromium(II)</td>
<td>chromous</td>
</tr>
<tr>
<td></td>
<td>Cr(^{3+})</td>
<td>chromium(III)</td>
<td>chromic</td>
</tr>
<tr>
<td>Cobalt</td>
<td>Co(^{2+})</td>
<td>cobalt(II)</td>
<td></td>
</tr>
<tr>
<td></td>
<td>Co(^{3+})</td>
<td>cobalt(III)</td>
<td></td>
</tr>
<tr>
<td>Copper</td>
<td>Cu(^{+})</td>
<td>copper(I)</td>
<td>cuprous</td>
</tr>
<tr>
<td></td>
<td>Cu(^{2+})</td>
<td>copper(II)</td>
<td>cupric</td>
</tr>
<tr>
<td>Iron</td>
<td>Fe(^{2+})</td>
<td>iron(II)</td>
<td>ferrous</td>
</tr>
<tr>
<td></td>
<td>Fe(^{3+})</td>
<td>iron(III)</td>
<td>ferric</td>
</tr>
<tr>
<td>Lead</td>
<td>Pb(^{2+})</td>
<td>lead(II)</td>
<td>plumbous</td>
</tr>
<tr>
<td></td>
<td>Pb(^{4+})</td>
<td>lead(IV)</td>
<td>plumbic</td>
</tr>
<tr>
<td>Mercury</td>
<td>Hg(^{2+})</td>
<td>mercury (I)</td>
<td>mercurous</td>
</tr>
<tr>
<td></td>
<td>Hg(^{2+})</td>
<td>mercury (II)</td>
<td>mercuric</td>
</tr>
<tr>
<td>Tin</td>
<td>Sn(^{2+})</td>
<td>tin(II)</td>
<td>stannous</td>
</tr>
<tr>
<td></td>
<td>Sn(^{4+})</td>
<td>tin(IV)</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:

\[ _\text{H}C_2\text{H}_3\text{O}_2 \] (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:

\[ \text{H}_2\text{S}(aq) = \text{hydrosulfuric acid; HI}(aq) = \text{hydroiodic acid; HN}_3(aq) = \text{hydrazoic acid} \]

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):

\[ \text{H}_2\text{S}(g) = \text{hydrogen sulfide; HI}(g) = \text{hydrogen iodide; HN}_3(l) = \text{hydrogen azide} \]
Oxoacid Nomenclature

If the acid contains oxygen ($HOX$ or $HXO_n$) 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 $XO_n^-$); so:

– **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*

$H_2SO_4(aq) \rightarrow$ anion is $SO_4^{2-}$ (sulfate) $\rightarrow$ acid name is *sulfuric acid*
$HClO(aq) \rightarrow$ anion is $ClO^-$ (hypochlorite) $\rightarrow$ acid name is *hypochlorous acid*

Acid Nomenclature: Practice

Try the following:

- $H_2C_2O_4$ $HIO_4$
- $H_2SO_3$ $H_2CO_3$
- $HCN$ $HF$
- $H_2CrO_4$ $HC_2H_3O_2$
- $H_3PO_4$ $HNO_3$
- $H_2Se$ $HNO_2$

*Give yourself five minutes*
Acid Nomenclature: Practice

Answers:

\[ H_2C_2O_4 \quad \text{oxalic acid} \]
\[ H_2SO_3 \quad \text{sulfurous acid} \]
\[ HCN \quad \text{hydrocyanic acid} \]
\[ H_2CrO_4 \quad \text{chromic acid} \]
\[ H_3PO_4 \quad \text{phosphoric acid} \]
\[ H_2Se \quad \text{hydroseLENic acid} \]

Acid Nomenclature: Practice

Answers:

\[ HIO_4 \quad \text{periodic acid} \]
\[ H_2CO_3 \quad \text{carbonic acid} \]
\[ HF \quad \text{hydrofluoric acid} \]
\[ HC_2H_3O_2 \quad \text{acetic acid} \]
\[ HNO_3 \quad \text{nitric acid} \]
\[ HNO_2 \quad \text{nitrous acid} \]
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

Sample Problem 2.2  Calculating the Mass of an Element in a Compound

PROBLEM: Pitchblende is the most commercially important compound of uranium. Analysis shows that 84.2 g of pitchblende contains 71.4 g of uranium, with oxygen as the only other element. How many grams of uranium can be obtained from 102 kg of pitchblende?

PLAN: The mass ratio of uranium/pitchblende is the same no matter the source. We can use the ratio to find the answer.

\[
\text{Mass (kg) of pitchblende} \times \text{mass ratio of uranium to pitchblende from analysis} \rightarrow \text{Mass (kg) of uranium}
\]

\[
1 \text{ kg} = 1000 \text{ g}
\]

\[
\text{Mass (g) of uranium}
\]
Calculating the Mass of an Element in a Compound

**Sample Problem 2.2**

**PROBLEM:** Silicon (Si) is essential to the computer industry as a major component of semiconductor chips. It has three naturally occurring isotopes: $^{28}$Si, $^{29}$Si, and $^{30}$Si. Determine the number of protons, neutrons, and electrons in each silicon isotope.

**PLAN:** We have to use the atomic number and atomic masses.

**SOLUTION:** The atomic number of silicon is 14. Therefore,

- $^{28}$Si has 14 protons, 14 electrons, and 14 neutrons ($28 - 14$)
- $^{29}$Si has 14 protons, 14 electrons, and 15 neutrons ($29 - 14$)
- $^{30}$Si has 14 protons, 14 electrons, and 16 neutrons ($30 - 14$)
Sample Problem 2.5  Calculating the Atomic Mass of an Element

**PROBLEM:** Silver (Ag: Z = 47) has 46 known isotopes, but only two occur naturally, $^{107}$Ag and $^{109}$Ag. Given the following mass spectrometric data, calculate the atomic mass of Ag:

<table>
<thead>
<tr>
<th>Isotope</th>
<th>Mass (amu)</th>
<th>Abundance (%)</th>
</tr>
</thead>
<tbody>
<tr>
<td>$^{107}$Ag</td>
<td>106.90509</td>
<td>51.84</td>
</tr>
<tr>
<td>$^{109}$Ag</td>
<td>108.90476</td>
<td>48.16</td>
</tr>
</tbody>
</table>

**PLAN:** We have to find the weighted average of the isotopic masses, so we multiply each isotopic mass by its fractional abundance and then sum those isotopic portions.

**SOLUTION:**

mass portion from $^{107}$Ag = 106.90509 amu x 0.5184 = 55.42 amu

mass portion from $^{109}$Ag = 108.90476 amu x 0.4816 = 52.45 amu

atomic mass of Ag = 55.42 amu + 52.45 amu = 107.87 amu
Sample Problem 2.6  Predicting the Ions an Element Forms

PROBLEM: What monatomic ions do the following elements form?

(a) Iodine (Z = 53)  (b) Calcium (Z = 20)  (c) Aluminum (Z = 13)

PLAN: Use Z to find the element. Find its relationship to the nearest noble gas. Elements occurring before the noble gas gain electrons and elements following lose electrons.

SOLUTION:

(a) Iodine is a nonmetal in Group 7A(17). It gains one electron to have the same number of electrons as 54Xe.

Ca^{2+}  Calcium is a metal in Group 2A(2). It loses two electrons to have the same number of electrons as 18Ar.

Al^{3+}  Aluminum is a metal in Group 3A(13). It loses three electrons to have the same number of electrons as 10Ne.

Sample Problem 2.15  Calculating the Molecular Mass of a Compound

PROBLEM: Using the data in the periodic table, calculate each molecular (or formula) mass:

(a) Tetraphosphorus trisulfide  (b) Ammonium nitrate

PLAN: Write the formula and then multiply the number of atoms (in the subscript) by the respective atomic masses. Add the masses for the compound.

SOLUTION:

(a) P_4S_3  
\[ \text{molecular mass} = (4 \times \text{atomic mass of P}) + (3 \times \text{atomic mass of S}) \]
\[ = (4 \times 30.97 \text{ amu}) + (3 \times 32.07 \text{ amu}) \]
\[ = 220.09 \text{ amu} \]

(b) NH_4NO_3  
\[ \text{molecular mass} = (2 \times \text{atomic mass of N}) + (4 \times \text{atomic mass of H}) + (3 \times \text{atomic mass of O}) \]
\[ = (2 \times 14.01 \text{ amu}) + (4 \times 1.008 \text{ amu}) + (3 \times 16.00 \text{ amu}) \]
\[ = 80.05 \text{ amu} \]