Chapter 2
Atoms, Molecules, and Ions

Jim Geiger
Cem 151
Atomic Theory of Matter

The theory of atoms:
Original to the Greeks
Leucippus, Democritus and Lucretius
(Aristotle thought they were nuts)
John Dalton (1805-1808)
Revived the idea and made it science by measuring the atomic weights of 21 elements.
That’s the key thing because then you can see how elements combine.
Dalton’s Postulates

Each element is composed of extremely small particles called atoms.

Tiny balls make up the world
Dalton’s Postulates

All atoms of a given element are identical to one another in mass and other properties, but the atoms of one element are different from the atoms of all other elements.
Dalton’s Postulates

Atoms of an element are not changed into atoms of a different element by chemical reactions; atoms are neither created nor destroyed in chemical reactions. (As far as Dalton knew, they couldn’t be changed at all).

Red O’s stay Os and blue N’s stay N’s.
Dalton’s Postulates

Compounds are formed when atoms of more than one element combine; a given compound always has the same relative number and kind of atoms.

Chemistry happens when the balls rearrange

[Diagram showing the reaction of H and N to form NH₃ (ammonia)]
Law of Constant Composition

Joseph Proust (1754–1826)

- Also known as the law of definite proportions.
- The elemental composition of a pure substance never varies.
- The relative amounts of each element in a compound doesn’t vary.

ammonia always has 3 H and 1 N.
Law of Conservation of Mass

The total mass of substances present at the end of a chemical process is the same as the mass of substances present before the process took place.

\[ 3H_2 + N_2 \rightarrow 2NH_3 \]

ammonia

The atoms on the right all appear on the left.
Streams of negatively charged particles were found to emanate from cathode tubes.

J. J. Thompson (1897).

Maybe atoms weren’t completely indivisible after all.
Thompson measured the charge/mass ratio of the electron to be $1.76 \times 10^8$ coulombs/g.

How? by manipulating the magnetic and electrical fields and observing the change in the beam position on a fluorescent screen.
Millikan Oil Drop Experiment

measured charge of electron
Univ. Chicago (1909).

How?
Vary the electric field (E) until the drops stop.
Vary the charge (q) on the drop with more X-rays. Get a multiple of $1.6 \times 10^{-19}$ Coulombs. The charge of 1 electron.

$Eq = mg$
You set E, measure mass of drop (m) & know g. Find q.
Radioactivity:

- The spontaneous emission of radiation by an atom.
- First observed by Henri Becquerel.
- Also studied by Marie and Pierre Curie.

“rays” not particles
particles of some sort.

Stuff comes out of atoms, “subatomic particles”
Radioactivity

• Three types of radiation were discovered by Ernest Rutherford:
  - \( \alpha \) particles, attracted to negative electrode, so they have a positive charge, much more mass than negative stuff (turn out to be He nuclei)
  - \( \beta \) particles, attracted to positive electrode, so they have a negative charge, 1000s of times less massive (turn out to be electrons coming from nucleus).
The Atom, circa 1900:

- “Plum pudding” model, put forward by Thompson.
- Positive sphere of matter with negative electrons imbedded in it.
- Most of the volume = positive stuff because most of the mass is positive.
- Expectation: density more or less uniform throughout.
Discovery of the Nucleus

Ernest Rutherford shot $\alpha$ particles at a thin sheet of gold foil and observed the pattern of scatter of the particles.
The Nuclear Atom

Virtually all the particles went straight through
Most of the atom essentially empty
A few particles deflected, some straight back.
A very small part of the atom is very dense, impenetrable.
The mass must be concentrated.
The Nuclear Atom

• Rutherford postulated a very small, dense nucleus with the negative electrons around the outside of the atom.
• Most of the volume of the atom is empty space.
Other Subatomic Particles

- Protons were discovered by Rutherford in 1919. Have the positive charge in the atom.
- Neutrons were discovered by James Chadwick in 1932. Have mass like proton, but no charge.
Subatomic Particles

- Protons and electrons are the only particles that have a charge.
- Protons and neutrons have essentially the same mass.
- The mass of an electron is so small we ignore it.

<table>
<thead>
<tr>
<th>Particle</th>
<th>Charge</th>
<th>Mass (amu)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Proton</td>
<td>Positive (1+)</td>
<td>1.0073</td>
</tr>
<tr>
<td>Neutron</td>
<td>None (neutral)</td>
<td>1.0087</td>
</tr>
<tr>
<td>Electron</td>
<td>Negative (1−)</td>
<td>$5.486 \times 10^{-4}$</td>
</tr>
</tbody>
</table>
Symbols of Elements

Mass number (number of protons plus neutrons)

Atomic number (number of protons or electrons)

$^{12}_6\text{C}$ — Symbol of element

Elements are symbolized by one or two letters.
Atomic Number

Mass number (number of protons plus neutrons)

Atomic number (number of protons or electrons)

All atoms of the same element have the same number of protons:
The atomic number (Z)
Atomic Mass

The mass of an atom in atomic mass units (amu) is approximately the total number of protons and neutrons in the atom.
Isotopes:

- Elements are defined by the number of protons.
- Atoms of the same element with different masses.
- Isotopes have different numbers of neutrons.

\[
\begin{align*}
\text{Neutrons} & \quad 5 & \quad 6 & \quad 7 & \quad 8 \\
\text{Isotopes} & \quad ^{11}\text{C}_{6} & \quad ^{12}\text{C}_{6} & \quad ^{13}\text{C}_{6} & \quad ^{14}\text{C}_{6} \\
\end{align*}
\]
Atomic and molecular masses can be measured with great accuracy with a mass spectrometer. Heavier ion turns less in the magnetic field (magnetic moments of ions similar).
Average Mass

• Because in the real world all the elements exist as mixtures of isotopes.
• And we measure many many atoms at a time
  "Natural abundance"
• Average mass is calculated from the isotopes of an element weighted by their relative abundances.
Average mass, example

<table>
<thead>
<tr>
<th>Isotope</th>
<th>abundance</th>
<th>Atomic mass</th>
</tr>
</thead>
<tbody>
<tr>
<td>$^{24}\text{Mg}$</td>
<td>78.99%</td>
<td>23.98504 amu</td>
</tr>
<tr>
<td>$^{25}\text{Mg}$</td>
<td>10.00%</td>
<td>24.98584 amu</td>
</tr>
<tr>
<td>$^{26}\text{Mg}$</td>
<td>11.01%</td>
<td>25.98259 amu</td>
</tr>
</tbody>
</table>

Given the above data, what is the average molecular mass of magnesium (Mg)?

\[
.7899(23.98504)+0.1000(24.98584)+0.1101(25.98259)=18.95 + 2.499 + 2.861 = 24.31
\]
### Periodic Table:

- **A systematic catalog of elements.**
- **Elements are arranged in order of atomic number.**

![Periodic Table Image]
When one looks at the chemical properties of elements, one notices a repeating pattern of reactivities.
### Periodic Table

- The rows on the periodic chart are periods.
- Columns are groups.
- Elements in the same group have similar chemical properties.
These five groups are known by their names. You gotta know these very well (except for the chalcogens, its far less common).

<table>
<thead>
<tr>
<th>Group</th>
<th>Name</th>
<th>Elements</th>
</tr>
</thead>
<tbody>
<tr>
<td>1A</td>
<td>Alkali metals</td>
<td>Li, Na, K, Rb, Cs, Fr</td>
</tr>
<tr>
<td>2A</td>
<td>Alkaline earth metals</td>
<td>Be, Mg, Ca, Sr, Ba, Ra</td>
</tr>
<tr>
<td>6A</td>
<td>Chalcogens</td>
<td>O, S, Se, Te, Po</td>
</tr>
<tr>
<td>7A</td>
<td>Halogens</td>
<td>F, Cl, Br, I, At</td>
</tr>
<tr>
<td>8A</td>
<td>Noble gases (or rare gases)</td>
<td>He, Ne, Ar, Kr, Xe, Rn</td>
</tr>
</tbody>
</table>
Nonmetals are on the upper right-hand corner of the periodic table (with the exception of H).
Metalloids border the stair-step line (with the exception of Al and Po).
Atoms, Molecules, and Ions

Periodic Table

Metals are on the left side of the chart.
• Elements required for living organisms (pretty much all organisms).
• Red, most abundant
• blue, next most abundant
• Green, trace amounts.
Chemical Formulas

The subscript to the right of the symbol of an element tells the number of atoms of that element in one molecule of the compound.
Molecular Compounds

Molecular compounds are composed of molecules and almost always contain only nonmetals.
Diatomic Molecules

These seven elements occur naturally as molecules containing two atoms.
You should know these guys.
Types of Formulas

• Empirical formulas give the lowest whole-number ratio of atoms of each element in a compound.

• Molecular formulas give the exact number of atoms of each element in a compound.

Example: ethane:

Empirical formula: \( \text{CH}_3 \)
Molecular formula: \( \text{C}_2\text{H}_6 \)
Types of Formulas

Structural formulas show the order in which atoms are bonded.

Perspective drawings also show the three-dimensional array of atoms in a compound.
Ions

When atoms lose or gain electrons, they become ions. Often they lose or gain electrons to have the same number of electrons as the nearest noble gas.

- Cations are positive and are formed by elements on the left side of the periodic chart.
- Anions are negative and are formed by elements on the right side of the periodic chart.
**Mono-atomic ions**

- Metals usually become cations (+)
- Nonmetals usually become anions (-)
Ionic compounds

- A metal will give up electrons to a nonmetal forming a cation (+) (the metal), and an anion (-) (the nonmetal).

\[ \text{Na} + \text{Cl} \rightarrow \text{Na}^+ + \text{Cl}^- \rightarrow \text{NaCl} \]
\[ \text{Mg} + 2\text{Cl} \rightarrow \text{Mg}^{2+} + 2\text{Cl}^- \rightarrow \text{MgCl}_2 \]

Note, everybody gains or loses electrons to be like the nearest noble gas.

Compounds are always electrically neutral!!
Writing Formulas

Because compounds are electrically neutral, one can determine the formula of a compound this way:

- The charge on the cation becomes the subscript on the anion.
- The charge on the anion becomes the subscript on the cation.
- If these subscripts are not in the lowest whole-number ratio, divide them by the greatest common factor.

\[
\begin{align*}
\text{Mg}^{2+} & \quad \text{O}^{2-} \quad \rightarrow \quad \text{MgO} \quad \text{Not Mg}_2\text{O}_2
\end{align*}
\]
# Common Cations

<table>
<thead>
<tr>
<th>Charge</th>
<th>Formula</th>
<th>Name</th>
<th>Formula</th>
<th>Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>1+</td>
<td>*H⁺</td>
<td>Hydrogen ion</td>
<td>*NH₄⁺</td>
<td>Ammonium ion</td>
</tr>
<tr>
<td></td>
<td>*Li⁺</td>
<td>Lithium ion</td>
<td>Cu⁺</td>
<td>Copper(I) or cuprous ion</td>
</tr>
<tr>
<td></td>
<td>*Na⁺</td>
<td>Sodium ion</td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td>*K⁺</td>
<td>Potassium ion</td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td>*Cs⁺</td>
<td>Cesium ion</td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td>Ag⁺</td>
<td>Silver ion</td>
<td></td>
<td></td>
</tr>
<tr>
<td>2+</td>
<td>*Mg²⁺</td>
<td>Magnesium ion</td>
<td>Co²⁺</td>
<td>Cobalt(II) or cobaltous ion</td>
</tr>
<tr>
<td></td>
<td>*Ca²⁺</td>
<td>Calcium ion</td>
<td>Cu²⁺</td>
<td>Copper(II) or cupric ion</td>
</tr>
<tr>
<td></td>
<td>*Sr²⁺</td>
<td>Strontium ion</td>
<td>Fe²⁺</td>
<td>Iron(II) or ferrous ion</td>
</tr>
<tr>
<td></td>
<td>*Ba²⁺</td>
<td>Barium ion</td>
<td>Mn²⁺</td>
<td>Manganese(II) or manganous ion</td>
</tr>
<tr>
<td></td>
<td>Zn²⁺</td>
<td>Zinc ion</td>
<td>Hg²⁺</td>
<td>Mercury(I) or mercurous ion</td>
</tr>
<tr>
<td></td>
<td>Cd²⁺</td>
<td>Cadmium ion</td>
<td></td>
<td>Mercury(II) or mercuric ion</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>Ni²⁺</td>
<td>Nickel(II) or nickleous ion</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>Pb²⁺</td>
<td>Lead(II) or plumbous ion</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>Sn²⁺</td>
<td>Tin(II) or stannous ion</td>
</tr>
<tr>
<td>3+</td>
<td>*Al³⁺</td>
<td>Aluminum ion</td>
<td>Cr³⁺</td>
<td>Chromium(III) or chromic ion</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>Fe³⁺</td>
<td>Iron(III) or ferric ion</td>
</tr>
</tbody>
</table>

*The most common ions are in boldface.

*You should know these.*
# Common Anions

<table>
<thead>
<tr>
<th>Charge</th>
<th>Formula</th>
<th>Name</th>
<th>Charge</th>
<th>Formula</th>
<th>Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>1−</td>
<td>*H(^{-})</td>
<td>Hydride ion</td>
<td></td>
<td>*C(_2)H(_3)O(_2)(^{-})</td>
<td>Acetate ion</td>
</tr>
<tr>
<td></td>
<td>*F(^{-})</td>
<td>Fluoride ion</td>
<td></td>
<td>*ClO(_3)(^{-})</td>
<td>Chlorate ion</td>
</tr>
<tr>
<td></td>
<td>*Cl(^{-})</td>
<td>Chloride ion</td>
<td></td>
<td>*ClO(_4)(^{-})</td>
<td>Perchlorate ion</td>
</tr>
<tr>
<td></td>
<td>*Br(^{-})</td>
<td>Bromide ion</td>
<td></td>
<td>*NO(_3)(^{-})</td>
<td>Nitrate ion</td>
</tr>
<tr>
<td></td>
<td>*I(^{-})</td>
<td>Iodide ion</td>
<td></td>
<td>*MnO(_4)(^{-})</td>
<td>Permanganate ion</td>
</tr>
<tr>
<td></td>
<td>*CN(^{-})</td>
<td>Cyanide ion</td>
<td></td>
<td>*ClO(_2)</td>
<td>Chlorite</td>
</tr>
<tr>
<td></td>
<td>*OH(^{-})</td>
<td>Hydroxide ion</td>
<td></td>
<td>*ClO</td>
<td>Hypochlorite</td>
</tr>
<tr>
<td>2−</td>
<td>O(^2)(^{-})</td>
<td>Oxide ion</td>
<td></td>
<td>*CO(_3)(^2)(^{-})</td>
<td>Carbonate ion</td>
</tr>
<tr>
<td></td>
<td>O(_2)(^2)(^{-})</td>
<td>Peroxide ion</td>
<td></td>
<td>*CrO(_4)(^2)(^{-})</td>
<td>Chromate ion</td>
</tr>
<tr>
<td></td>
<td>S(^2)(^{-})</td>
<td>Sulfide ion</td>
<td></td>
<td>*Cr(_2)O(_7)(^2)(^{-})</td>
<td>Dichromate ion</td>
</tr>
<tr>
<td>3−</td>
<td>N(^3)(^{-})</td>
<td>Nitride ion</td>
<td></td>
<td>*PO(_4)(^3)(^{-})</td>
<td>Phosphate ion</td>
</tr>
</tbody>
</table>

*The most common ions are in boldface.

*You should know these.
More polyatomic anions (the “ites”)

<table>
<thead>
<tr>
<th>Ion</th>
<th>Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>SCN⁻</td>
<td>Thiocyanate</td>
</tr>
<tr>
<td>NO₂⁻</td>
<td>Nitrite</td>
</tr>
<tr>
<td>HSO₃⁻</td>
<td>bisulfite</td>
</tr>
<tr>
<td>HSO₄⁻</td>
<td>bisulfate</td>
</tr>
<tr>
<td>HPO₄²⁻</td>
<td>Hydrogen phosphate</td>
</tr>
<tr>
<td>H₂PO₄⁻</td>
<td>Dihydrogen phosphate</td>
</tr>
<tr>
<td>ClO⁻</td>
<td>hypochlorite</td>
</tr>
<tr>
<td>ClO₂⁻</td>
<td>chlorite</td>
</tr>
</tbody>
</table>
Patterns in Oxyanion Nomenclature

• When there are **only two** oxyanions involving the same element:
  
  ➢ The one with fewer oxygens ends in *-ite*
    
    - $\text{NO}_2^-$: nitrite; $\text{SO}_3^{2-}$: sulfite
  
  ➢ The one with more oxygens ends in *-ate*
    
    - $\text{NO}_3^-$: nitrate; $\text{SO}_4^{2-}$: sulfate
Patterns in Oxyanion Nomenclature
When there are more than two:

- The one with the second fewest oxygens ends in \(-\text{ite}\)
  - \(\text{ClO}_2^–\): chlorite
- The one with the second most oxygens ends in \(-\text{ate}\)
  - \(\text{ClO}_3^–\): chlorate
- The one with the fewest oxygens has the prefix \(\text{hypo-}\) and ends in \(-\text{ite}\)
  - \(\text{ClO}^–\): hypochlorite
- The one with the most oxygens has the prefix \(\text{per-}\) and ends in \(-\text{ate}\)
  - \(\text{ClO}_4^–\): perchlorate
Inorganic Nomenclature

• Write the name of the cation.
• If the anion is an element, change its ending to -ide; if the anion is a polyatomic ion, simply write the name of the polyatomic ion.
• If the cation can have more than one possible charge, write the charge as a Roman numeral in parentheses.
Examples
naming inorganic compounds

• Write the name of the cation.
• If the anion is an element, change its ending to "-ide"; if the anion is a polyatomic ion, simply write the name of the polyatomic ion.
• If the cation can have more than one possible charge, write the charge as a Roman numeral in parentheses.

NaCl  
NH₄NO₃  
Fe(SO₄)  
KCN  
RbOH  
LiC₂H₃O₂  
NaClO₃  
NaClO₄  
K₂CrO₄  
NaH  
sodium chloride  
ammonium nitrate  
Iron(II) sulfate  
potassium cyanide  
Rubidium hydroxide  
lithium acetate  
sodium chlorate  
sodium perchlorate  
potassium chromate  
Sodium hydride
Examples
naming inorganic compounds

- Write the name of the cation.
- If the anion is an element, change its ending to -ide; if the anion is a polyatomic ion, simply write the name of the polyatomic ion.
- If the cation can have more than one possible charge, write the charge as a Roman numeral in parentheses.

NaCl  sodium chloride
potassium permanganate  KMnO₄
Calcium carbonate  CaCO₃
Calcium bicarbonate  Ca(HCO₃)₂
ammonium dichromate  NH₄(Cr₂O₇)
potassium phosphate  (K)₃PO₃
Lithium oxide  Li₂O  (O²⁻ is the anion)
sodium peroxide  Na₂O₂  (O₂²⁻ is the anion)
Calcium sulfide  CaS
Hydrogen

- H can be cation or anion
- H\(^-\) hydride
- H\(^+\) (the cation of an inorganic compound) makes an acid, naming different.
Acid Nomenclature

- If the anion in the acid ends in *-ide*, change the ending to *-ic acid* and add the prefix *hydro-*:
  - HCl: hydrochloric acid
  - HBr: hydrobromic acid
  - HI: hydroiodic acid
Acid Nomenclature

- If the anion in the acid ends in -ite, change the ending to -ous acid:
  - HClO: hypochlorous acid
  - HClO₂: chlorous acid
Acid Nomenclature

- If the anion in the acid ends in -ate, change the ending to -ic acid:
  - HClO₃: chloric acid
  - HClO₄: perchloric acid
Nomenclature of Binary Compounds

- The less electronegative atom is usually listed first.
- A prefix is used to denote the number of atoms of each element in the compound (\textit{mono} is not used on the first element listed, however.)

<table>
<thead>
<tr>
<th>Prefix</th>
<th>Meaning</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mono-</td>
<td>1</td>
</tr>
<tr>
<td>Di-</td>
<td>2</td>
</tr>
<tr>
<td>Tri-</td>
<td>3</td>
</tr>
<tr>
<td>Tetra-</td>
<td>4</td>
</tr>
<tr>
<td>Penta-</td>
<td>5</td>
</tr>
<tr>
<td>Hexa-</td>
<td>6</td>
</tr>
<tr>
<td>Hepta-</td>
<td>7</td>
</tr>
<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 of Binary Compounds

- The ending on the more electronegative element is changed to \(-ide\).

<table>
<thead>
<tr>
<th>Prefix</th>
<th>Meaning</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mono-</td>
<td>1</td>
</tr>
<tr>
<td>Di-</td>
<td>2</td>
</tr>
<tr>
<td>Tri-</td>
<td>3</td>
</tr>
<tr>
<td>Tetra-</td>
<td>4</td>
</tr>
<tr>
<td>Penta-</td>
<td>5</td>
</tr>
<tr>
<td>Hexa-</td>
<td>6</td>
</tr>
<tr>
<td>Hepta-</td>
<td>7</td>
</tr>
<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>

- \(\text{CO}_2\): carbon dioxide
- \(\text{CCl}_4\): carbon tetrachloride
Nomenclature of Binary Compounds

If the prefix ends with a or o and the name of the element begins with a vowel, the two successive vowels are often merged into one:

\[ \text{N}_2\text{O}_5: \text{dinitrogen pentoxide} \]

not: \( \text{dinitrogen pentaoxide} \)
Atoms,
Molecules,
and Ions
Ionic Bonds

Ionic compounds (such as NaCl) are generally formed between metals and nonmetals.