Chapter 2. Atoms, Molecules, and Ions

2.1 The Atomic Theory of Matter: The Early History

As stated in Chapter 1, chemistry's primary goal is to explain the macroscopic behavior of matter in terms of the atomic scale structures and interactions. This chapter provides brief historical sketch of the major steps that led our modern atomic theory that accomplishes this goal.

Ancient Theories

The ancient Greek philosopher Leucippus and his pupil Democritus (470-380 BC.) were the first to speculate that matter is made of indivisible particles called atomos (a = "not" + tomos = "to cut"). Their theory was not widely accepted in their time. Most Greek philosophers, including Aristotle (384-322 BC.), believed that matter was infinitely divisible (i.e., continuous). Aristotle further speculated that all material derived their properties from the various proportions of just four elements: earth, air, fire, and water. Their view prevailed for 2000 years.

While the Greeks can be given credit for trying to provide rational explanations for the physical world, they never adopted the modern scientific approach of using experiments to "test" scientific models or theories. Their ideas did not amount to scientific theories in the modern sense. Rather, they were mere speculations aimed at fitting observations to preconceived notions about how the world should be.

From pre-historic times curiosity and practical needs have led enterprising and curious humans to manipulate their material environment, beginning with bone and stone tools, fire, and native metals. The Medieval alchemists (500 AD-1500 AD) combined such practical interests with mystical strivings to search for a "philosopher's stone," a substance that could transform base metals (copper, iron, zinc, etc.) into gold. Alchemical "theories" of transmutation seem strange by today's standards. Ultimately, none of their ideas could be supported by experiment.

Although the Alchemists' primary goal was never achieved, their work resulted in many practical techniques and knowledge that supported the development of modern chemistry later on.

The Dawn of the Modern Atomic Theory

In 1661, Robert Boyle defined an element as a substance that cannot be broken down into simpler substances by chemical operations (e.g., heating). This represented the modern and eminently practical idea of an operational definition. It was based on practical procedures and facts and not on preconceived notions or mystical beliefs about the ultimate nature of matter.

In 1789, Antoine Lavoisier published the first chemistry textbook that listed 33 elements. While a few of these elements were later found to be compounds, the notion of fundamental substances with unique and definite properties was firmly established.
Lavoisier’s experiments utilized very sensitive balances to measure and compare the mass of reactants and products in chemical reactions suggested the idea that matter is neither created nor destroyed. That is, he promoted the Law of Conservation of Mass:

Matter is neither created nor destroyed in any physical or chemical process. The total mass of the reacting substances equals the mass of the products.

Lavoisier’s use of the mass balance placed chemistry on a firm quantitative footing and led to other important chemical laws. In 1799, Joseph Louis Proust discovered that copper carbonate, whether obtained from natural sources or synthesized in the laboratory, always contained copper, oxygen, and carbon in the same proportions by mass, (i.e., 5.3 parts Copper to 4.0 parts Oxygen to 1.0 parts Carbon). Similar results with other substances led to the formulation of the Law of Definite Proportions:

Compounds always contain elements in certain definite proportions by mass.

**Dalton’s Atomic Theory**

In 1803, John Dalton published the first modern atomic theory to explain the unique identity of elements, what happens in a chemical reaction, the Law of Conservation of Mass, the Law of Definite Proportions, and the Law of Multiple Proportions.

Dalton’s Atomic Theory consisted of five postulates. These postulates along with the explanations their provided are:

1. All elements are made up of tiny, indivisible particles called atoms. Atoms can neither be created nor destroyed during chemical reactions.  
   *This postulate explains the conservation of mass.*

2. All atoms of a given element are identical, but the atoms of one element differ from the atoms of another.  
   *Along with the first postulate, this explains why elements cannot be broken down into simpler substances.*

3. Atoms of different elements form compounds by joining in fixed, small, whole number ratios to form molecules.  
   *This postulate explains the law of definite proportions. If compounds are made of molecules with fixed number ratios of atoms then they must also have fixed mass ratios of elements.*

4. A chemical reaction involves a change not in the atoms themselves, but in the way the atoms are combined to form compounds.  
   *This postulate explains chemical reactions and emphasizes the fundamental nature of atoms.*

Dalton proposed that the atoms of each element are distinguished by their mass. He assigned the hypothetical hydrogen atom, the lightest, a relative mass of 1. He and others developed a table of atomic weights of the other elements from experimentally measured mass proportions of compounds and guessing the number ratio of atoms in the compound (i.e., the compound’s formula).
Example: The mass ratio of oxygen to hydrogen in water is 8:1. The mass ratio of carbon to oxygen in carbon dioxide is 3:8. Determine the atomic masses of oxygen and carbon relative to hydrogen (which is arbitrarily assigned the value of 1 atomic mass unit) given that the formula of water is $H_2O$ (2 atoms $H$ to 1 atom $O$) and the formula of carbon dioxide is $CO_2$ (1 atom $C$ to 2 atoms $O$).

You can probably figure out the required atomic masses in (please try!). Oxygen would have an atomic mass of 16 and carbon would then be 12. However, the following mathematical model can help you quickly and confidently calculate them is very useful for more complicated cases.

Let $\frac{m_A}{m_B}$ be the measured definite mass ratio of element $A$ to element $B$ in a compound of $A$ and $B$.

Let $\frac{N_A}{N_B}$ be the hypothetical (guessed) number ratio of $A$ atoms to $B$ atoms in a molecule of the compound.

Let $M_A$ and $M_B$ be the atomic masses of atoms of elements $A$ and $B$ relative to hydrogen, which is arbitrarily assigned the value 1 atomic mass unit. Then, Dalton’s Atomic Theory explains the Law of Definite through the expression

$$\frac{m_A}{m_B} = \frac{N_A M_A}{N_B M_B}$$

where the left side expresses the empirical measurements that go into the Law of Definition Proportions and the right side expresses the theoretical justification. Applying this expression to oxygen and hydrogen in water, assuming the atom number ratio of hydrogen to water is $\frac{N_H}{N_O} = 2/1$ leads to

$$\frac{m_O}{m_H} = \frac{N_O M_O}{N_H M_H} \Rightarrow M_O = \frac{N_H}{N_O} \cdot \frac{m_O}{m_H} \cdot M_H = \left(\frac{2}{1}\right) \cdot \left(\frac{8}{1}\right) \cdot (1) = 16$$

And applying this formula for carbon in carbon dioxide given that $M_O = 16$ gives

$$\frac{m_C}{m_O} = \frac{N_C M_C}{N_O M_O} \Rightarrow M_C = \frac{N_O}{N_C} \cdot \frac{m_C}{m_O} \cdot M_O = \left(\frac{2}{1}\right) \cdot \left(\frac{3}{8}\right) \cdot (16) = 12$$

Although Dalton’s theory successfully explained many chemical laws, it did not prove that atoms existed. Atoms were considered hypothetical particles. Strong evidence for the actual existence of atoms was not available until the early 1900’s.

The Law of Multiple Proportions

Soon after proposing his atomic theory John Dalton announced the Law of Multiple Proportions and was able to employ his theory to provide a convincing explanation.

When two elements form more than one compound, the ratio of the mass ratio of first element to the second element in one compound to the mass ratio of first element to the second in the other compound is a simple whole number ratio.

Example: Nitrogen and oxygen form three compounds that show the law of multiple proportions

<table>
<thead>
<tr>
<th>Compound number</th>
<th>Compound name</th>
<th>$\frac{\text{mass } O}{\text{mass } N}$</th>
<th>$\frac{\left(\frac{\text{mass } O}{\text{mass } N}\right)}{\frac{\text{mass } O}{\text{mass } N}}$</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>Nitrous oxide</td>
<td>0.571</td>
<td>1</td>
</tr>
<tr>
<td>2</td>
<td>Nitric oxide</td>
<td>1.142</td>
<td>2</td>
</tr>
<tr>
<td>3</td>
<td>Nitrogen dioxide</td>
<td>2.284</td>
<td>4</td>
</tr>
</tbody>
</table>

2-3
The Law of Multiple Proportions and its explanation by Dalton’s atomic theory can be mathematically modeled by simply extending the simple Definite Proportions model outlined above. The expression for the Law is given on the left side and the following expression and the theory on the right side.

\[
\frac{\text{mass } A}{\text{mass } B}_{\text{cmpd } 1} = \frac{N_A \cdot M_A}{N_B \cdot M_B}_{\text{cmpd } 1} = \frac{N_A}{N_B}_{\text{cmpd } 1}
\]

\[
\frac{\text{mass } A}{\text{mass } B}_{\text{cmpd } 2} = \frac{N_A \cdot M_A}{N_B \cdot M_A}_{\text{cmpd } 2} = \frac{N_A \cdot M_A}{N_B}_{\text{cmpd } 2}
\]

All of the symbols have the same meaning as before. That the ratio of mass ratios turns out (empirically) to be a ratio of simple integers (the Law of Multiple Proportions) follows directly from a model of compounds having as their smallest unit molecules that have simple integer number ratios of atoms.

Example: If nitrous oxide molecules (see table above) have the formula \(N_2O\) (each molecule contains two atoms of nitrogen and one atom of oxygen) then what must be the formula of nitric oxide?

Applying the above expression leads to:

\[
\frac{\text{mass } O}{\text{mass } N}_{\text{cmpd } 1} = \frac{N_O}{N_N}_{\text{cmpd } 1} \Rightarrow \frac{1}{1.142} = \frac{1}{2} = \frac{N_O}{N_N}_{\text{cmpd } 2}
\]

Therefore, the simplest possible formula is \(NO\). But it also may be true that the actual molecules are \(N_2O_2\) or \(N_2O_2\), we cannot tel.

Problems with Dalton’s Theory

By the early 20th century experimental evidence had arisen that showed that Dalton's Theory was incomplete and only approximate. For example

1. Nonstoichiometric compounds having a range of mass ratios of elements were found.
2. Not all of the atoms of a given element had precisely the same mass. These were called isotopes.
3. Atoms are not indestructible and have a subatomic structure. The discovery of the “radioactive decay” of some elements led to the development of nuclear fission and fusion.

These "chinks" in the armor of Dalton's Theory led to the development of the modern theory of the atom during the late 19th century and early 20th century. Some of the major steps of this development will be given here, with the rest taken up in Chapter 6.

Still, great progress in chemistry followed Dalton’s program. Dozens of elements were discovered by the mid 19th century and the relative atomic masses were “weighed” by relating the relative masses of elements in compounds to the hypothetical number ratio of atoms in the compound.

2.2. The Discovery of Atomic Structure

In 1879, William Crookes discovered cathode rays. These were beams of particles that were emitted from a highly charged negative metal electrodes embedded inside a partially evacuated tube. As the cathode rays passed through the gas they made it glow. When the rays struck a phosphorescent screen, they would cause it to glow also. They could not, however, pass through the glass walls of the tube.
In 1895 Wilhelm Roentgen discovered X-rays. These were later determined to be high-energy light emitted when cathode rays struck a metal target. X-rays traveled many meters through air outside the tube. Their ability to pass through flesh, but not as well bone, led to X-ray photography.

In 1896 Antoine Bequerel, Madame Curie, and Pierre Curie observed that uranium spontaneously emitted X-rays without cathode ray bombardment. They call it the process radioactivity. Their discovery suggested that "something" was coming out of the atom. In other words, atoms have "insides" or parts.

In 1897 J. J. Thomson demonstrated that cathode rays are negatively charged particles by repelling them with a strong negative charge and by deflecting them sideways with a magnet. He names them electrons. Thomson also measures the charge to mass ratio of the electron \( \frac{q}{m} \). Thomson concluded that electrons are subatomic particles and a common building block of all matter. The following experimental results supported this conclusion.

1. All electrons are identical no matter what metal the cathode material.
2. Since electrons have negative charge the other parts of the atom must be positively charged.
3. The cathode does not lose mass. If cathode rays were whole atoms the cathode would eventually show a loss of mass.

In 1904 J. J. Thomson developed his "plum pudding" model of atom. It depicted electrons imbedded like raisins in a continuous positively charged pudding-like mass. Thomson's model was more sophisticated than this sounds. He worked out the precise positions the electrons must occupy within the positive charge to be stable.

In 1909 Robert Millikan determined the charge on an electron by performing a Newton’s Second law analysis of the terminal velocities of tiny charged oil drops falling between electrically charged plates. This charge was hypothesized to be the elementary (or smallest) unit of charge \( 1.6 \times 10^{-19} \) Coloumb). They don’t get smaller!

Millikan’s determination of the electron charge allowed the mass of the electron to be calculated from Thomson’s \( \frac{q}{m} \) ratio. That is:

---

1 A negatively charged particle moving across a flat-face of the north pole of a magnet will be deflected sideways. This important effect led to the modern means of measuring atomic masses, as we shall soon see.

2 We will see precisely how this experiment works in Experiment 14. It utilizes several concepts that we learned in Honors Physics.
\[
m_e = \frac{e}{e/m} = \frac{-1.6 \times 10^{-19} \text{Coulomb}}{-1.76 \times 10^8 \text{Coulomb/gram}} = 9.11 \times 10^{-28} \text{g}
\]

The mass of the electron was \(\frac{1}{1827}\) the mass of a hydrogen atom, which supported Thomson’s theory that the electron is a subatomic particle.

In 1907 Ernest Rutherford identified three subatomic particles emitted from radioactive elements. He named them alpha, beta, and gamma particles after the first three letters of the Greek alphabet. These particles were later shown to the positive nucleus of a helium atom, electrons, and gamma light rays, respectively. This was further evidence that atoms have parts; there exists a subatomic structure.

<table>
<thead>
<tr>
<th>Particle Name</th>
<th>Symbol</th>
<th>Mass (amu)</th>
<th>Charge</th>
<th>Later Identified as</th>
</tr>
</thead>
<tbody>
<tr>
<td>alpha</td>
<td>(\alpha)</td>
<td>4</td>
<td>+2</td>
<td>Helium nucleus</td>
</tr>
<tr>
<td>beta</td>
<td>(\beta)</td>
<td>1/837</td>
<td>-1</td>
<td>electron</td>
</tr>
<tr>
<td>gamma</td>
<td>(\gamma)</td>
<td>0</td>
<td>0</td>
<td>High energy light</td>
</tr>
</tbody>
</table>

In 1911 Rutherford and Marsden performed an experiment in which they bombarded a very thin gold foil with positively charged alpha particles emitted from radioactive uranium (see Figure at top of next page). While Thomson’s “plum pudding” model predicted that the particles would fly straight through without slowing down, they found that a very small percentage were deflected through very large angles. The rare deflections suggested that the vast majority of an atom’s mass was concentrated in a very tiny positively charged nucleus. Most positive alpha-particles pass straight through because they miss any nucleus by wide margins. Only a few pass close enough to a nuclei to suffer a large deflection. A statistical analysis of the scattering allowed Rutherford to deduce the nuclear diameter to be about 1/10,000 the diameter of the atom! Atoms were mostly empty space. This evidence suggested a planetary model of the atom in which negatively charged electrons orbited the tiny, but massive, positively charged nucleus, held in orbit by the attractive electrical forces. To make the atom neutral, the number of orbiting electrons must equal the nuclear charge.

\[\text{\textsuperscript{3}\textsuperscript{3}}\] The beta particle is therefore identical to a cathode ray. They are both electrons. The distinction is the source. Cathode rays are produced by strong electric fields that “pull” them from the outer “orbit” of atoms. Beta particles originate from inside the atomic nucleus during a radioactive decay process.
2.3. A Preview of the "Modern" View of Atomic Structure

While Rutherford's planetary model still lingers in the minds of nonscientists, it actually came under immediate attack because it contradicted the well-established theory of electromagnetic radiation. This theory predicted that an orbiting electron would continuously convert its kinetic energy into light energy. As a result it should spiral into the nucleus. Still, there was no doubt that an atom was comprised of a very small, massive, positive nucleus and electrons that were held close to it by electrical attractions.

Neil's Bohr, a Danish physicist and student of Rutherford, was the first to propose an alternate model. His model was supported by the discoveries and theories of Albert Einstein and others in the early 20th century. The major threads of Bohr's contribution to atomic theory, as well its subsequent development into the most modern form of atomic theory, called quantum mechanics, will be studied up in Chapter 6.

For now it is worth describing just a few of the features of modern atomic theory. Remember, however, that a more complete picture will be needed to explain the chemical properties and transformations of matter, which, after all, is the ultimate goal of Chemistry.

Subatomic Particles: Electrons, Protons, and Neutrons

Atoms are not indivisible as Dalton hypothesized. They are made up protons, neutrons, and electrons which have the properties given in the table below

<table>
<thead>
<tr>
<th>Particle Name</th>
<th>Symbol</th>
<th>Charge (e)</th>
<th>Approx mass (amu)</th>
<th>Precise Mass (g)</th>
<th>Size (m)</th>
</tr>
</thead>
<tbody>
<tr>
<td>electron</td>
<td>e⁻</td>
<td>-1</td>
<td>1/1837</td>
<td>9.10953×10⁻²⁸</td>
<td>“point”</td>
</tr>
<tr>
<td>proton</td>
<td>p⁺</td>
<td>+1</td>
<td>1</td>
<td>1.67265×10⁻²⁴</td>
<td>&lt;1×10⁻¹⁵</td>
</tr>
<tr>
<td>neutron</td>
<td>n</td>
<td>0</td>
<td>1</td>
<td>1.67495×10⁻²⁴</td>
<td>&lt;1×10⁻¹⁵</td>
</tr>
</tbody>
</table>

1 amu = 1.6606×10⁻²⁴ g
1 e⁻ = 1.6×10⁻¹⁹ C

The Structure and Size of the Atoms

The protons and neutrons of an atom are located in a very tiny region called the nucleus. The diameter of atomic nuclei is on the order of 10⁻¹⁵ m. The electrons "swarm" around the nucleus in regions called orbitals. The electrons determine the overall size of atoms that are on the order of 10⁻¹⁰ m. That is, atomic size is determined by this region occupied by the electrons. To give an idea of this scale difference between the nucleus and the whole atom, if an atom was magnified to the size of a baseball stadium, the nucleus would be the size of a marble.

In the late 1980’s the invention of tunneling electron microscopes allowed images of atoms and dramatically confirm these scales. Mr. Kemer took the micrograph below during his sabbatical appointment as a Visiting Scholar at Amherst College in 2000. It shows neatly aligned molecules of docosyl-ether and docosyl-sulfide lying on a flat surface of graphite. The smallest white dots are hydrogen atoms. The brightest dots are sulfur atoms. The figure to the right shows models of these same molecules (the yellow center atom represents sulfur and the red center represents oxygen). The blurriness of the images arises in part to the motions of the molecules during the scan that produced the image. The image is about 10 nanometers (10 billionths of a meter) across.
Atomic Number and Mass Number

All atoms of a particular element have the same number of protons in their nuclei. This number is called its **atomic number** (Z). For instance, Hydrogen has 1 proton in its nucleus and therefore has the atomic number 1. *It is the atomic number that identifies the element.*

The **mass number** (A) is defined as the total number of protons and neutrons in the nucleus.

\[
\text{Mass number} = \text{atomic number} + \text{neutron number} \quad (A = Z + A)
\]

A neutral atom has equal numbers of electrons and protons.

**Example:** Fill in the blanks of the following table. Use a periodic table or other source for the symbols.

<table>
<thead>
<tr>
<th>Atomic Number</th>
<th>Mass Number</th>
<th>Number of Protons</th>
<th>Number of Neutrons</th>
<th>Number of Electrons*</th>
<th>Element Symbol</th>
<th>Element name</th>
</tr>
</thead>
<tbody>
<tr>
<td>7</td>
<td>7</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td>nitrogen</td>
</tr>
<tr>
<td>39</td>
<td>10</td>
<td></td>
<td></td>
<td>19</td>
<td></td>
<td></td>
</tr>
<tr>
<td>59</td>
<td>27</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

2.4. Atomic Weights

**Measuring The Masses of Atoms**

The mass of atoms and molecules is accomplished using a **mass spectrometer**. The basis of mass spectroscopy is the law of magnetic force on a moving charged particle. This law, called the Lorentz Force law, states that a charged particle will be deflected sideways into a circular path when it moves across the face of a magnet having its north pole pointing perpendicular to the particles velocity. The magnitude of the force is proportional to the particle charge and speed and the strength of the magnetic field.

\[
F_B = qvB \quad F = ma \quad qvB = m\frac{v^2}{R} \quad m = \frac{qBR}{v}
\]

A detailed diagram of a mass spectrometer is shown below. Note that the atoms are stripped of an electron to give them a single positive elementary charge. This loss of mass must be added in to give the correct mass of the atom. Note that the north pole of the magnet points down.
Isotopes

Two atoms of the same element must have the same number of protons. However, they may have different numbers of neutrons and therefore different mass numbers and masses. Atoms of a given element with different mass numbers are called isotopes (“same position”). Although isotopes of the same element differ in mass, they generally have very similar chemical properties. The symbol system used to designate an isotope is

\[ \frac{A}{Z} X \]

where \( X \) is the element symbol, \( Z \) is the atomic number, and \( A \) is the mass number. Often the atomic number is not written since it is implied by the element symbol.

Isotopes are also expressed with the element name followed by a hyphen and the mass number.

Example: Two isotopes of chlorine: \( ^{35}Cl \) (chlorine-35) and \( ^{37}Cl \) (chlorine-37)

Example: Two isotopes of Carbon: \( ^{12}C \) (carbon-12) and \( ^{14}C \) (carbon-14)

Example: Hydrogen has three isotopes. They are the only isotopes given specific names.

\( ^{1}H \) (Protium) \( ^{2}H \) (Deuterium) \( ^{3}H \) (Tritium)

Example: Complete the following table using the definitions of atomic number, mass number.

<table>
<thead>
<tr>
<th>isotope</th>
<th>symbol</th>
<th>Atomic number</th>
<th>Mass number</th>
<th>No. of protons</th>
<th>No. of neutrons</th>
</tr>
</thead>
<tbody>
<tr>
<td>Carbon-12</td>
<td>( ^{12}C )</td>
<td></td>
<td></td>
<td>6</td>
<td>6</td>
</tr>
<tr>
<td></td>
<td>( ^{82}Sr )</td>
<td></td>
<td>92</td>
<td>29</td>
<td>61</td>
</tr>
</tbody>
</table>

The Atomic Weight Scale and the Mole

The mass spectrometer allows the masses of atoms (minus an electron) to be measured in gram units. For instance, the \( ^{1}H \) isotope has a mass of \( 1.6735 \times 10^{-24} \) g and the \( ^{16}O \) isotope has a mass of \( 2.6560 \times 10^{-23} \) g. Because it would be cumbersome to continually express such small masses in grams, we instead use a unit called the atomic mass unit, or amu. The amu is defined in as 1/12 the mass of the \( ^{12}C \) isotope.

\[ 1 \text{ amu} = 1.66054 \times 10^{-24} \text{ g} \]

A proton has a mass of 1.0073 amu, a neutron 1.0087 amu, and an electron \( 5.486 \times 10^{-4} \) amu.

The mole is the number defined as the number of \( ^{12}C \) atoms in exactly 12.0 grams of \( ^{12}C \). This is also the number of amu units in 1.0 gram

\[ 1 \text{ mole} = 6.022 \times 10^{23} \]

The clever thing about this definition of a mole is that the mass of 1 mole of an element (or compound) is the number of grams numerically equal to its mass in amu. This will be made more clear in Chapter 3.

Average Atomic Masses of the Elements

The atomic weight of an element is a weighted average mass of its isotopes that reflects the relative abundances of each isotope as they occur naturally.
Example: A sample of chlorine contains 75.78% of $^{35} \text{Cl}$ and 24.22% of $^{37} \text{Cl}$ by mass. Calculate the atomic weight of chlorine given that the atomic mass of $^{35} \text{Cl}$ is 34.969 amu and $^{37} \text{Cl}$ is 36.966 amu.

\[
(0.7578)(34.969 \text{ amu}) + (0.2423)(36.966 \text{ amu}) = 35.45 \text{ amu}
\]

Example: A naturally occurring sample of Chlorine, which contains of $^{35} \text{Cl}$ and $^{37} \text{Cl}$, has an average atomic weight of 35.45 amu. What is the relative abundance of each isotope?

Let x be the fractional relative abundance of $^{35} \text{Cl}$. Approximate the atomic mass of these two isotopes to be 35 and 37.

\[(x)(35 \text{ amu}) + (1 - x)(37 \text{ amu}) = 35.45 \text{ amu}
\]

Solving for x leads to $x = 0.76$. Therefore,

\[
\% ^{35} \text{Cl} = (100\%) x = 76\%, \quad \% ^{37} \text{Cl} = (100\%)(1-x) = 24\%
\]

The relative abundances of isotopes are determined by mass spectroscopy. On a mass spectrograph the relative heights of the peaks is directly related to the mass percentages of each isotope.

Example: The mass spectrograph for chlorine on the top right figure of page 2-9 shows the two peaks for the isotopes of chlorine. The relative heights of the peaks give the relative mass % abundances. In this case the relative heights of the $^{35}\text{Cl}$ and $^{37}\text{Cl}$ peaks are 5.0 and 1.6, respectively. Therefore, naturally occurring chlorine contains $5/(5+1.6) = 76\% \text{ } ^{35}\text{Cl}$ and $1.6/(5+1.6) = 24\% \text{ } ^{37}\text{Cl}$ by mass.

2.5. The Periodic Table

In 1814, J. J. Berzelius invented modern element symbols. Today we know that each element is made up of a unique type of atom. 89 elements (atoms) occur naturally on the earth. 20 elements have been synthesized since 1937 by smashing smaller atoms (nuclei) together at nearly the speed of light in particle accelerators. Most synthesized atoms radioactively decay into stable elements in very short times

Abundant and Rare Elements

The matter of the known universe is 93% H, 7% He, and < 0.1% all else. The solar system is 85% H, 15% He, and < 0.1% all else. Strangely, more than 60% of the mass of the Universe is of unknown composition; it is simply called “dark matter.”

Periodic Table of the Elements

In 1869 Dimitri Mendeleev, a Russian Chemist, organized the known elements into a table in order of increasing atomic weights. He chose to start new horizontal rows (or periods) of the table in such a way that elements with similar chemical and physical properties lined up in vertical columns (or groups).
Mendeleev's ordering scheme resulted in some "empty" places that he took for undiscovered elements. Indeed, on this basis he predicted the existence of scandium (Sc), gallium (Ga) and germanium (Ge), all of which were discovered shortly afterward.

Modern atomic theory has determined that the true basis for the ordering the elements in the periodic table is increasing atomic number and that the vertical groups are comprised of elements whose atoms have similar arrangements on electrons in the atoms. In other words, modern atomic theory relates chemical properties to the electronic structure of atoms.

Elements are commonly classified as follows:

<table>
<thead>
<tr>
<th>Type</th>
<th>Physical properties</th>
</tr>
</thead>
<tbody>
<tr>
<td>Metals</td>
<td>Solid at room temperature (except Hg), good conductors of heat and electricity, shiny, ductile, have a wide range of hardness.</td>
</tr>
<tr>
<td>Nonmetals</td>
<td>May be solid, liquid, and gas at room temperature, are poor conductors of heat and electricity, nonmetal solids have dull surfaces and are brittle.</td>
</tr>
<tr>
<td>Metalloids</td>
<td>have properties are intermediate to metals and nonmetals. Some have unique electrical properties (semiconductors)</td>
</tr>
<tr>
<td>Diatomic Elements</td>
<td>These exist naturally as diatomic molecules. They include: $H_2$, $N_2$, and $O_2$ plus</td>
</tr>
<tr>
<td>Noble</td>
<td>These do not ordinarily form compounds with other elements. They are all</td>
</tr>
<tr>
<td>Gases</td>
<td>monatomic gases at room temperature. They include $He$, $Ne$, $Ar$, $Kr$, $Xe$, $Rn$</td>
</tr>
</tbody>
</table>

The group designations include:
- The Group IA metals are called the Alkali metals
- The Group IIA metals are called the Alkaline Earth metals
- The Group 1B-8B metals (center portion of the periodic table) are called the Transition Metals
- The Group VIIA nonmetals are called the Halogens
- The section of elements 57-70 and 89-102 are called the Inner Transition Metals
- The other Groups are simply referred to by their Group number or by the element at the top.
Modern periodic tables include a wealth of physical and chemical information for each element including the average atomic weight. It is important to review three things about the atomic weight listed in periodic tables:

1. It represents the weighted average of all the naturally occurring isotopes of that element.
2. If interpreted as amu units, it gives the average mass of a single atom.
3. If interpreted as gram units, it gives the mass of 1 mole of atoms of that element.

2.6. Molecules and Molecular Compounds

The atom is the smallest representative sample of an element. However, only the noble gas elements (He, Ne, Ar, Rn) are normally found in nature as isolated atoms. Most matter is composed of molecules or ions, both of which are formed from atoms. We examine molecules here and ions in the next section.

A molecule is an assembly of two or more atoms tightly bound together. The resultant "package" of atoms behaves in many ways as a single, distinct object. We will discuss the forces that hold the atoms together (the chemical bonds) in Chapters 8 and 9.

Molecules and Chemical Formulas

Many elements are found in nature in molecular form; that is, two or more of the same type of atom are bound together. For example, the oxygen normally found in air consists of molecules that contain two oxygen atoms. We represent this molecular form of oxygen by the chemical formula $O_2$ (read "oh two"). The subscript in the formula tells us that two oxygen atoms are present in each molecule. Any molecule that is made up of two atoms is said to be a diatomic molecule. The other diatomic elements include hydrogen ($H_2$), nitrogen ($N_2$), fluorine ($F_2$), chlorine ($Cl_2$), bromine ($Br_2$), and iodine ($I_2$).

Oxygen also exists in another molecular form known as ozone. Molecules of ozone consist of three oxygen atoms, so its chemical formula is $O_3$. Even though "normal" oxygen ($O_2$) and ozone are both composed only of oxygen atoms, they exhibit very different chemical and physical properties. For example, $O_2$ is essential for life, but $O_3$ is toxic; $O_2$ is odorless, whereas $O_3$ has a sharp, pungent smell. Elements that exist in more than one form are called allotropes.

Substances that are composed of molecules made up or two or more types of atoms are called molecular compounds. For example, water is a molecular compound comprised of molecules made of two hydrogen atoms and one oxygen atom. It is represented by the chemical formula $H_2O$. Lack of a subscript on the $O$ indicates one atom of $O$ per water molecule. Another molecular compound composed of these same elements is hydrogen peroxide, $H_2O_2$. The properties of these two compounds are very different. Several molecules of other molecular compounds are shown in Figure 2.16. Notice how the composition of each compound is given by its chemical formula. Notice also that these substances are composed only of nonmetallic elements. Molecular substances contain only nonmetals.

![Figure 2.16. Molecular models of some simple molecules.](image)
Molecular and Empirical Formulas

Chemical formulas that indicate the actual numbers and types of atoms in a molecule are called **molecular formulas** (the formulas in Figure 2.16 are molecular formulas). Chemical formulas that give only the relative number of atoms of each type in a molecule are called **empirical formulas**. The subscripts in an empirical formula always give the smallest whole-number ratios of the atoms.

**Example:** The molecular formula for hydrogen peroxide is $H_2O_2$; its empirical formula is $HO$.

**Example:** The molecular formula for ethylene is $C_2H_4$; its empirical formula is $CH_2$.

For many substances, the molecular formula and the empirical formula are identical, as in the case of water, $H_2O$.

Whenever we know the molecular formula of a compound, we can determine its empirical formula. However, the converse is not true; if we know the empirical formula of a substance, we can't determine its molecular formula unless we know more information. So why do chemists concern themselves with empirical formulas? As we will see in Chapter 3, common methods of analyzing substances lead to the empirical formula only. Once the empirical formula is known, additional experiments can give the information needed to convert the empirical formula to the molecular one. In addition, there are substances such ionic compounds that do not exist as molecules. For these substances only empirical formulas are meaningful.

**Example:** Write the empirical formulas for the following molecules:

(a) glucose (a.k.a. blood sugar or dextrose) whose molecular formula is $C_6H_{12}O_6$ : $CH_2O$
(b) nitrous oxide (a.k.a laughing gas) whose molecular formula is $N_2O$: Same as empirical formula

Picturing Molecules

The molecular formula of a substance summarizes its composition but does not show how its atoms join together. For many molecules it is possible to draw **structural formulas** that show the connectivity of the atoms. For example the structural formulas of water, hydrogen peroxide, and methane ($CH_4$) are:

![Structural formulas](image)

Usually a structural formula does not depict the actual geometry of the molecule. That is, it cannot show the actual angles at which atoms are joined together in three-dimensions. A structural formula can be written, however, as a perspective drawing to give some sense of three-dimensional shape, as shown in Figure 2.17.
Scientists also rely on various other types of models to help them visualize molecules. Ball-and-stick models show atoms as spheres and the bonds as sticks, and they give a good sense of the angles at which the atoms are attached to one another within the molecule. Often, all atoms are represented by balls of the same size. In other cases, however, the relative sizes of the balls reflect the relative sizes of the atoms. Sometimes the chemical symbols of the elements are superimposed on the balls, but often the atoms are identified simply by color.

Space-filling models give a semi-accurate depiction of what the molecule would look like if it were scaled up in size. These models show the relative sizes of the atoms, but the angles between atoms, which help define their molecular geometry, are often more difficult to see than in ball-and-stick models. As in ball-and-stick models, the identities of the atoms are indicated by their colors, but they may also be labeled with the element’s symbol.

You might wonder how chemists determine all of the details of these structures. This is a fascinating question that we will address later in the course. If you do not care to wait, try doing a little research yourself. Google is a convenient place to start.

2.7. Ions and Ionic Compounds

The nucleus of an atom is unchanged by ordinary chemical processes, but atoms can readily gain or lose electrons. If electrons are removed or added to a neutral atom, a charged particle called an ion is formed. An ion with a positive charge is called a cation (pronounced CAT-ion); a negatively charged ion is called an anion (AN-ion).

For example, the neutral sodium atom, which has 11 protons and 11 electrons, easily loses one electron. The resulting cation has 11 protons and 10 electrons, and hence has a net charge of 1+. The net charge on an ion is represented by a superscript; +, 2+, and 3+ mean a net charge resulting from the loss of one, two, or three electrons, respectively. The superscripts -, 2-, and 3- represent net charges resulting from the gain of one, two, or three electrons, respectively. The formation of the Na⁺ ion from a Na atom is shown schematically below:

![Na atom losing an electron to form Na⁺ ion](image)

Chlorine, with 17 protons and 17 electrons, gains an electron in chemical reactions, producing the Cl⁻ ion:

![Cl atom gaining an electron to form Cl⁻ ion](image)

(In general, metal atoms tend to lose electrons; nonmetal atoms tend to gain electrons)

In addition to simple ions such as Na⁺ and Cl⁻, there are polyatomic ions such as $NO_3^-$ (nitrato ion) and $SO_4^{2-}$ (sulfato ion). These ions consist of atoms joined as in a molecule, but they have a net positive or negative charge. We will consider further examples of polyatomic ions in Section 2.8. The chemical properties of ions are greatly different from those of the atoms from which they are derived.)
Determining Ionic Charges

Many atoms gain or lose electrons so as to end up with the same number of electrons as the noble gas closest to them in the periodic table. The members of the noble gas family are chemically very non-reactive and form very few compounds. We might deduce that this is because their electron arrangements are somehow complete and, therefore, energetically stable. The other elements can obtain these same stable electron arrangements by losing or gaining electrons. For example, the loss of one electron from an atom of sodium leaves it with the same number of electrons as the neutral neon atom (atomic number 10). Similarly, when a chlorine atom gains an electron, it ends up with 18, the same as argon (atomic number 18). We will use this simple rule to predict the formation of ions.

Example: Predict the charges expected for the most stable ions of barium (Ba) and oxygen (O).

From the periodic table, barium has atomic number 56. The nearest noble gas is xenon, atomic number 54. Barium can most readily obtain the stable arrangement of 54 electrons by losing two of its electrons, forming the $\text{Ba}^{2+}$ cation. Oxygen has atomic number 8. The nearest noble gas is neon, atomic number 10. Oxygen can most readily obtain this stable electron arrangement by gaining two electrons, thereby forming an anion of 2- charge, $\text{O}^{2-}$. In general, metals tend to lose electrons to form positive cations and nonmetals tend to gain electrons to form negative anions.

The periodic table can be used to determine the charges of ions. The table below shows how the charges of these ions relate in a simple way to their positions in the table. On the left side of the table, we see that the group 1A elements (the alkali metals) form 1+ ions, and the group 2A elements (the alkaline earths) form 2+ ions. On the other side of the table, the group 7A elements (the halogens) form 1- ions, and the group 6A elements form 2- ions. The transition metals do not lend themselves to such simple rules and many can form more than one charge.

<table>
<thead>
<tr>
<th>1A</th>
<th>2A</th>
<th>3A</th>
<th>4A</th>
<th>5A</th>
<th>6A</th>
<th>7A</th>
<th>8A</th>
</tr>
</thead>
<tbody>
<tr>
<td>H+</td>
<td>Li+</td>
<td>Na+</td>
<td>Mg2+</td>
<td>Al3+</td>
<td>S2-</td>
<td>Cl-</td>
<td>F-</td>
</tr>
<tr>
<td>K+</td>
<td>Ca2+</td>
<td>Rb+</td>
<td>Sr2+</td>
<td>Cs+</td>
<td>Ba2+</td>
<td>Tranision metals</td>
<td></td>
</tr>
</tbody>
</table>

Ionic Compounds

A great deal of chemical activity involves the transfer of electrons between substances. Ions form when one or more electrons transfer from one neutral atom to another. In Figure 2.19 below we see that when elemental sodium is allowed to react with elemental chlorine, an electron transfers from a neutral sodium atom to a neutral chlorine atom. We are left with a $\text{Na}^{+}$ ion and a $\text{Cl}^{-}$ ion. But that is not the end of the matter. Because objects of opposite charge attract, the $\text{Na}^{+}$ and $\text{Cl}^{-}$ ions bind together to form $\text{NaCl}$, an ionic compound. An ionic compound is a compound that contains positively charged ions and negatively charged ions.

We can often tell whether a compound is ionic or molecular (consisting of molecules) from its composition. Ionic compounds are generally combinations of metals and nonmetals, as in $\text{NaCl}$. In contrast, molecular compounds are generally composed of nonmetals only, as in $\text{H}_2\text{O}$, $\text{C}_6\text{H}_12\text{O}_6$, and $\text{NO}_2$. 
Ionic compounds are also identified by the fact that they do not conduct electricity in solid form, but do conduct electricity when molten (liquid) or dissolved in water. This behavior is good evidence that ionic compounds are made of charged ions. When in the molten state or dissolved in water, the charged ions are mobile and free to respond to an external electric field; in other words, conduct electricity. For this reason ionic compounds are called electrolytes.

![Image](image.png)

**FIGURE 2.19** (a) The transfer of an electron from a neutral Na atom to a neutral Cl atom leads to the formation of an Na\(^+\) ion and a Cl\(^-\) ion. (b) Arrangement of these ions in solid sodium chloride, NaCl, pictured at the right.

**Example:** Which of the following compounds would you expect to be ionic: Na\(_2\)O, CaCl\(_2\), SF\(_4\)?

**Answer:** Na\(_2\)O and CaCl\(_2\) are ionic compounds since they are made of a metal and non-metal. SF\(_4\) is a molecular compound since S and F are both non-metals.

The ions in an ionic compound are arranged in an extended three-dimensional structure called a lattice. The arrangement of Na\(^+\) and Cl\(^-\) ions in NaCl is shown in Figure 2.19. Note how each positive ion is surrounded by negative ion nearest neighbors and vice versa. This extended, well-ordered, three-dimensional structure distinguishes ionic compounds from molecular compounds. The basic unit of a molecular compound is a molecule - a self-contained neutral group of atoms linked together in a particular order (see Figure 2-16). Ionic crystals do not have smallest units per se. They are built up from these extended three-dimensional arrangements of ions held together by electrostatic forces. How many positive and negative ions must get together before you can say there is a piece of an ionic solid? That’s hard to say. It is certainly many times more than a few of each ion. The tiniest crystal just visible under a microscope would have trillions. On the hand, a single molecule constitutes a piece of a molecular compound.

Because we cannot identify ionic compounds with molecules, the formulas we use to represent them communicate only the relative numbers of positive and negative ions. For example, when we represent magnesium chloride by the formula MgCl\(_2\) we are NOT referring to a molecule containing one Mg and 2 Cl atoms. This formula means that the ratio of Mg\(^{2+}\) ions to Cl\(^-\) ions is 1:2. In other words, formulas for ionic compounds are empirical formulas.

It is a simple matter to write the empirical formula for an ionic compound if we know the charges of the ions of which it is composed. Chemical compounds are always electrically neutral. Consequently, the ions in an ionic compound always occur in such a ratio that the total positive charge is equal to the total negative charge. Thus, there is one Na\(^+\) to one Cl\(^-\) giving NaCl, one Ba\(^{2+}\) to two Cl\(^-\) giving BaCl\(_2\), and so forth. As you consider these and other examples, you will see that if the charges on the cation and anion are equal, the subscript on each ion will be 1 (and left unwritten). If the charges are not equal, the charge on one ion (without its sign) will become the subscript on the other ion:
Example: What are the empirical formulas of the compounds formed by (a) \( Al^{3+} \) and \( Cl^- \) ions; (b) \( Al^{3+} \) and \( O^{2-} \) ions; (c) \( Mg^{2+} \) and \( NO_3^- \) ions?

SOLUTION (a) \( AlCl_3 \) (b) \( Al_2O_3 \) (c) \( Mg(NO_3)_2 \). Note that in this last case, the formula for the \( NO_3^- \) must be enclosed in parentheses so that it is clear that the subscript 2 applies to all the atoms of that ion. Otherwise it might look like there were 32 oxygen atoms! \( Mg(NO_3)_2 \)

2.8. Naming Inorganic Compounds (Chemical Nomenclature)

The astronomical number of compounds that exist makes it essential that a system of chemical nomenclature be used. The International Union of Pure and Applied Chemistry (IUPAC) has developed such a system for assigning names and formulas to compounds. In this section you will learn this system as it applies to inorganic compounds.

A. Naming Methods for Monatomic Ions

1. Cations of representative metal elements having just one charge (or oxidation) state:

   \[
   \text{element name } \text{ion} \hspace{1cm} \text{Example: } Mg^{2+} \text{ is named } \text{the "magnesium ion."}
   \]

2. Anions of representative non metal elements having just one charge state:

   \[
   \text{element root-} \text{ide ion} \hspace{1cm} \text{Example: } Cl^- \text{ is named } \text{the "chloride ion."}
   \]

   \[
   P^{3-} \text{ is named } \text{the "phosphide" ion}
   \]

3. Cations with more than one possible charge states:

Several transition metals can form cations of different charge. Examples include copper \( (Cu^+ \text{ and } Cu^{2+}) \) and iron \( (Fe^{2+} \text{ and } Fe^{3+}) \). There are two methods for naming ions that exist in more than one charge state:

- The Stock System: A Roman numeral is put in parentheses after the element name to indicate the particular value of the cation charge.

  Examples: \( Cu^+ \text{ = copper (I) ion (read "copper one ion")} \)
  \( Cu^{2+} \text{ = copper (II) ion (read "copper two ion")} \)
  \( Fe^{2+} \text{ = iron (II) ion (read "iron two ion")} \)
  \( Fe^{3+} \text{ = iron (III) ion (read "iron three ion")} \)

- Classical system: The suffixes -ous and -ic are added to the element root name. The -ous is used for the lower possible charge state while the -ic is used for the higher charge.

  Examples: \( Cu^+ \text{ = cuprous ion; } Cu^{2+} \text{ = cupric ion} \)
  \( Fe^{2+} \text{ = ferrous ion; } Fe^{3+} \text{ = ferric ion} \)

In this course we will stick to the modern Stock system, that is, we will use the Roman numerals.
B. Naming Polyatomic Ions

Polyatomic ions are tightly bound groups of atoms that behave as a unit and carry a charge. They are "charged molecules." Common examples include the nitrate ion \( \text{NO}_3^- \), the sulfate ion \( \text{SO}_4^{2-} \), and the phosphate ion \( \text{PO}_4^{3-} \). The only common polyatomic cation is the ammonium ion \( \text{NH}_4^+ \). The names of common polyatomic anions are given Table 2.5 at the end of the chapter.

C. Names and Formulas of Binary Ionic Compounds

A binary ionic compound is made up of a monatomic cation and a monatomic anion.

1. Writing Formulas of Binary Ionic Compounds

Consider the ionic compound formed between \( \text{Mg} \) and \( \text{Cl} \).

- Mg atoms lose two electrons to become cations with a +2 charge. This is written as \( \text{Mg}^{2+} \).
- Cl atoms gain 1 electron each to become anions with a -1 charge. This is written as \( \text{Cl}^- \).
- Because ionic compounds are electrically neutral, there must be two \( \text{Cl}^- \) for each \( \text{Mg}^{2+} \) in the compound.
- Therefore the formula for this compound is \( \text{MgCl}_2 \).

In general, the ionic compound formed from the metal ion \( M^{x+} \) and a nonmetal ion \( A^{-y} \) is given by

\[
M_yA_x
\]

**Example:** Sodium oxide: The ions are \( \text{Na}^+ \), \( O^{2-} \) \( \Rightarrow \) \( \text{Na}_2\text{O} \)

**Example:** Iron (III) oxide: The ions are \( \text{Fe}^{3+} \) and \( O^{2-} \) \( \Rightarrow \) \( \text{Fe}_2\text{O}_3 \)

**Example:** Magnesium nitride: The ions are \( \text{Mg}^{2+} \) and \( N^{3-} \) \( \Rightarrow \) \( \text{Mg}_3\text{N}_2 \)

2. Naming Binary Ionic Compounds

Binary ionic compounds are named by writing the cation name followed by the anion name (-ide ending).

**Example:** \( \text{NaCl} \), sodium chloride
\( \text{AlN} \), aluminum nitride
\( \text{Na}_3\text{P} \), sodium phosphide

When a cation has more than one possible charge state the particular cation charge that it exists in must be indicated by a Roman numeral in parentheses or by the proper cation suffix (-ous or -ic)

**Example:** \( \text{CuO} \), copper (II) oxide or cupric oxide
\( \text{Cu}_2\text{O} \), copper (I) oxide or cuprous oxide

The anion charge along with the condition of electrical neutrality determines the cation charge.

**Example:** \( \text{SnO}_2 \), must be tin (IV) oxide since the oxygen anion always has a -2 charge.

Roman numerals are used on an "as needed" basis only. That is, they are only used when the cation is one that can have more than one charge.
D. Names and Formulas of Ionic Compounds Containing Polyatomic Ions

1. Formulas of Ionic Compounds with Polyatomic Ions

The procedure for writing the formulas of ionic compounds having polyatomic ions is similar to that used for binary ionic compounds.

Example: Calcium nitrate: \( \text{Ca}^{2+}, \text{NO}_3^{-1} \Rightarrow \text{Ca(NO}_3)_2 \)

Example: Ammonium phosphate: \( \text{NH}_4^+, \text{PO}_4^{3-} \Rightarrow (\text{NH}_4^+)\text{PO}_4 \)

Note that parentheses must be placed around the polyatomic ion when there is more than one of them in the formula. This is necessary to make the subscripts clear. Parentheses are never placed around monatomic ions.

2. Naming Ionic Compounds Having Polyatomic Ions

Simply name the ions, cation first and anion second.

Example: \( \text{NaC}_2\text{H}_3\text{O}_2 \), sodium acetate

Example: \( \text{Cu(C}_2\text{H}_3\text{O}_2)_2 \), copper (II) acetate

Note the necessity of using Roman numeral after any cation that is one of those that can have more than one charge.

E. Binary Molecular Compounds

Binary molecular compounds are composed of two nonmetallic elements. Prefixes are used to describe the number of each atom in molecules.

1. Naming Binary Molecular Compounds

Since molecular compounds have nothing to do with ionic charges, the formulas cannot be predicted and prefixes are needed to show how many atoms of a given element are in a given molecule.

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

To name a binary molecular compound

1. Recognize the compound as being binary molecular. It must contain no metal atoms
2. Add prefixes to each element to indicate the number of atoms of each element.
3. The second element is written with an -ide ending
4. The "o" at the end of the prefix mono- is dropped when the element name begins with a vowel.
5. The mono- prefix is omitted entirely when there is a single atom of the first element.
Examples:

$N_2O$: dinitrogen monoxide (note that the "o" is dropped from mono- before "oxide."

$PCl_3$: phosphorous trichloride

$AlCl_3$: aluminum chloride (WATCH OUT, this is an ionic compound so prefixes are not used!)

$CO_2$: carbon dioxide

$CO$: carbon monoxide (Note: drop the “o” on mono when the element name starts with a vowel. You don’t want to write monoxide)

Two important facts should be kept in mind when dealing with molecular compounds

• They are not made of ions so ionic charges cannot be used in determine or imply the formula.

• Nonmetallic elements can often combine in more than one way. For example, there are many molecules that have $C$ and $H$ each having different numbers of each element. That is why prefixes are needed

2. Writing Formulas of Binary Molecular Compounds.

The prefixes give the numbers of atoms and therefore the subscripts to use.

Write the element symbols with the numbers of atoms of each as subscripts (1 is understood)

Examples:

Carbon disulfide: $CS_2$  
Dichlorine pentoxide: $Cl_2O_5$

F. Nomenclature of Acids

Acids are molecular compounds that contain a hydrogen atom that breaks away from the rest of its molecule when it dissolves in water. When not dissolved in water they are neutral molecules. When dissolved, they break up into an $H^+$ ion and a negative anion. For this reason, when dissolved in water, acids can be thought of as ionic compounds with a hydrogen ion $H^+$ playing the role of the metal cation.

Examples:

In its pure gaseous state, hydrogen chloride is a molecular compound and is symbolically represented as $HCl(g)$. When it dissolved in water hydrogen chloride dissociates into $H^+(aq)$ and $Cl^-(aq)$ ions. In this dissolved state $HCl(aq)$ is called hydrochloric acid.

In its pure gaseous state $H_2S$ (g) is called hydrogen sulfide. When dissolved in water $H_2S$ (aq) is called hydrosulfuric acid

Note that to indicate that a compound is dissolved in water the symbol (aq) is placed after the compound symbol. This stands for aqueous solution (dissolved in water).
1. **Naming Acids**

   The general formula for an acid is:

   \[ H_a A_b (aq) \]

   where \( H \) is hydrogen, \( A \) is an anion, and the symbol \( (aq) \) indicates that it is dissolved in water.

   The following are the **rules for naming acids**.

1. When the anion is derived from a single atom, write: **hydro-anion stem-ic acid**

   **Examples:**
   - \( HBr (aq) \): \( Br^- \) = bromide ion \( \Rightarrow \) hydrobromic acid
   - \( HI (aq) \): \( I^- \) = iodide ion \( \Rightarrow \) hydroiodic acid
   - \( H_2S (aq) \): \( S^{2-} \) = sulfide ion \( \Rightarrow \) hydrosulfic acid

2. When the anion name ends in **-ite**, write: **anion stem-ous acid**

   **Examples:**
   - \( H_2SO_3(aq) \): \( SO_3^{2-} \) = sulfite ion \( \Rightarrow \) sulfurous acid
   - \( HClO(aq) \): \( ClO^- \) = hypochlorite ion \( \Rightarrow \) hypochlorous acid

3. When the anion ends in **-ate**, write **anion stem-ic acid**

   **Examples:**
   - \( HNO_3(aq) \): \( NO_3^- \) = nitrate ion \( \Rightarrow \) nitric acid
   - \( H_3PO_4 (aq) \): \( PO_4^{3-} \) = phosphate ion \( \Rightarrow \) phosphoric acid

2. **Writing Acid Formulas:** Use the above rules in reverse
G. Nomenclature of Hydrates

Many ionic compounds incorporate a specific number of water molecules per formula unit into their crystal structures. These water molecules form part of the crystal and do not make the salt appear “wet.” Ionic compounds that contain water in this manner are called hydrates and the water that is incorporated into the crystal is called water of hydration.

The water of hydration can usually be removed by heating at temperatures near 100 °C. An ionic compound that has had its water of hydration removed is called an anhydrous salt.

The formula of a hydrate consists of the formula of the ionic compound followed by "\( \cdot x \ \text{H}_2\text{O} \)” where \( x \) is the number of water molecules attached to each formula unit of ionic compound.

The name of the hydrate is given by the ionic compound name followed by prefix-hydrate where the prefix is that corresponding to the \( x \).

**Examples:** Copper (II) sulfate pentahydrate: \( \text{CuSO}_4 \cdot 5\text{H}_2\text{O} \)

The images above show the hydrate copper (II) sulfate on the left. In the center is the anhydrous copper (II) sulfate attained after gentle heating. The right photograph shows a drop of water added to the anhydrous copper (II) sulfate which results in the return to the hydrated form.

The “dehydration” of copper (II) sulfate pentahydrate may be written as a chemical equation:

\[
\text{CuSO}_4 \cdot 5\text{H}_2\text{O}(s) \rightarrow \text{heat} \quad \text{CuSO}_4(s) + 5\text{H}_2\text{O}(l)
\]

\( \text{hydrate} \quad \text{(anhydrous salt)} \quad \text{(water)} \)

Another example of a hydrated salt is barium chloride dihydrate, \( \text{BaCl}_2 \cdot 2\text{H}_2\text{O} \).
Table of Common Ions

<table>
<thead>
<tr>
<th>Cations</th>
<th></th>
<th>Anions</th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>Name</td>
<td>Symbol</td>
<td>Hydrated color</td>
<td>Name</td>
<td>Symbol</td>
</tr>
<tr>
<td>Aluminum</td>
<td>$\text{Al}^{3+}$</td>
<td>-</td>
<td>Acetate</td>
<td>$\text{CH}_3\text{COO}^-$</td>
</tr>
<tr>
<td>Ammonium</td>
<td>$\text{NH}_4^+$</td>
<td>-</td>
<td>Arsenide</td>
<td>$\text{As}^{3-}$</td>
</tr>
<tr>
<td>Barium</td>
<td>$\text{Ba}^{2+}$</td>
<td>-</td>
<td>Bromide</td>
<td>$\text{Br}^-$</td>
</tr>
<tr>
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<td>$\text{Cd}^{2+}$</td>
<td>yellow</td>
<td>Bromate</td>
<td>$\text{BrO}_3^-$</td>
</tr>
<tr>
<td>Calcium</td>
<td>$\text{Ca}^{2+}$</td>
<td>-</td>
<td>Carbine</td>
<td>$\text{C}^4-$</td>
</tr>
<tr>
<td>Chromium (II)</td>
<td>$\text{Cr}^{2+}$</td>
<td>varies</td>
<td>Carbonate</td>
<td>$\text{CO}_3^{2-}$</td>
</tr>
<tr>
<td>Chromium (III)</td>
<td>$\text{Cr}^{3+}$</td>
<td>blue</td>
<td>Hydrogen carbonate</td>
<td>$\text{HCO}_3^-$</td>
</tr>
<tr>
<td>Cobalt (II)</td>
<td>$\text{Co}^{2+}$</td>
<td>light red</td>
<td>Chloride†</td>
<td>$\text{Cl}^-$</td>
</tr>
<tr>
<td>Cobalt (III)</td>
<td>$\text{Co}^{3+}$</td>
<td>-</td>
<td>Chlorate†</td>
<td>$\text{ClO}_3^-$</td>
</tr>
<tr>
<td>Copper (I)</td>
<td>$\text{Cu}^+$</td>
<td>-</td>
<td>Chlorite†</td>
<td>$\text{ClO}_2^-$</td>
</tr>
<tr>
<td>Copper (II)</td>
<td>$\text{Cu}^{2+}$</td>
<td>pale blue</td>
<td>Hypochlorite†</td>
<td>$\text{ClO}^-$</td>
</tr>
<tr>
<td>Hydrogen</td>
<td>$\text{H}^+$</td>
<td>-</td>
<td>Perchlorate†</td>
<td>$\text{ClO}_4^-$</td>
</tr>
<tr>
<td>Iron (II)</td>
<td>$\text{Fe}^{2+}$</td>
<td>pale green</td>
<td>Chromate</td>
<td>$\text{CrO}_4^{2-}$</td>
</tr>
<tr>
<td>Iron (III)</td>
<td>$\text{Fe}^{3+}$</td>
<td>brown</td>
<td>Dichromate</td>
<td>$\text{Cr}_2\text{O}_7^{2-}$</td>
</tr>
<tr>
<td>Lead (II)</td>
<td>$\text{Pb}^{2+}$</td>
<td>-</td>
<td>Cyanide</td>
<td>$\text{CN}^-$</td>
</tr>
<tr>
<td>Lead (IV)</td>
<td>$\text{Pb}^{4+}$</td>
<td>-</td>
<td>Thiocyanate</td>
<td>$\text{SCN}^-$</td>
</tr>
<tr>
<td>Lithium</td>
<td>$\text{Li}^+$</td>
<td>-</td>
<td>Fluoride</td>
<td>$\text{F}^-$</td>
</tr>
<tr>
<td>Magnesium</td>
<td>$\text{Mg}^{2+}$</td>
<td>-</td>
<td>Hydride</td>
<td>$\text{H}^-$</td>
</tr>
<tr>
<td>Manganese (II)</td>
<td>$\text{Mn}^{2+}$</td>
<td>light pink</td>
<td>Hydroxide</td>
<td>$\text{OH}^-$</td>
</tr>
<tr>
<td>Manganese (III)</td>
<td>$\text{Mn}^{3+}$</td>
<td>-</td>
<td>Nitride</td>
<td>$\text{N}^3-$</td>
</tr>
<tr>
<td>Mercury (I)</td>
<td>$\text{Hg}^{2+}$</td>
<td>-</td>
<td>Nitrate</td>
<td>$\text{NO}_3^-$</td>
</tr>
<tr>
<td>Mercury (II)</td>
<td>$\text{Hg}^{2+}$</td>
<td>-</td>
<td>Nitrite</td>
<td>$\text{NO}_2^-$</td>
</tr>
<tr>
<td>Nickel (II)</td>
<td>$\text{Ni}^{2+}$</td>
<td>green</td>
<td>Oxalate</td>
<td>$\text{C}_2\text{O}_4^{2-}$</td>
</tr>
<tr>
<td>Nickel (III)</td>
<td>$\text{Ni}^{3+}$</td>
<td>green</td>
<td>Oxide</td>
<td>$\text{O}^{2-}$</td>
</tr>
<tr>
<td>Potassium</td>
<td>$\text{K}^+$</td>
<td>-</td>
<td>Permanganate</td>
<td>$\text{MnO}_4^-$</td>
</tr>
<tr>
<td>Sodium</td>
<td>$\text{Na}^+$</td>
<td>-</td>
<td>Peroxide</td>
<td>$\text{O}_2^{2-}$</td>
</tr>
<tr>
<td>Tin (II)</td>
<td>$\text{Sn}^{2+}$</td>
<td>-</td>
<td>Phospide</td>
<td>$\text{P}^{3-}$</td>
</tr>
<tr>
<td>Tin (IV)</td>
<td>$\text{Sn}^{4+}$</td>
<td>-</td>
<td>Phosphate</td>
<td>$\text{PO}_4^{3-}$</td>
</tr>
<tr>
<td>Strontium</td>
<td>$\text{Sr}^{2+}$</td>
<td>-</td>
<td>Phosphate</td>
<td>$\text{PO}_3^{3-}$</td>
</tr>
<tr>
<td>Silver</td>
<td>$\text{Ag}^+$</td>
<td>-</td>
<td>Selenide</td>
<td>$\text{Se}^{2-}$</td>
</tr>
<tr>
<td>Zinc</td>
<td>$\text{Zn}^{2+}$</td>
<td>-</td>
<td>Sulfide</td>
<td>$\text{S}^{2-}$</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

† Ions with bromine and iodine in place of chlorine are named similarly