Chapter Objectives:

• Learn the development of the atomic theory.
• Understand the basic structure of the atom.
• Understand the structure of the periodic table.
• Learn how to write formulas and name ionic and binary molecular compounds.

Chapter 2
Atoms, Ions, and Molecules:
Matter Starts Here

Mr. Kevin A. Boudreaux
Angelo State University
CHEM 1411 General Chemistry
Chemistry: The Science in Context (Gilbert, 4th ed, 2015)
www.angelo.edu/faculty/kboudrea

You only arrive at the right answer after making all possible mistakes. The mistakes began with the Greeks.

The Road to the Atomic Theory

Nothing exists except atoms and empty space; everything else is opinion.
Democritus
The ancient Greek philosopher Democritus (c. 460 - 370 BC) reasoned that if you cut a lump of matter into smaller and smaller pieces, you would eventually cut it down to a particle which could not be subdivided any further. He called these particles **atoms** (from the Greek *atomos*, “uncutable”)

Aristotle (384-322 BC) believed that matter was continuous, and elaborated the idea that everything was composed of four elementary substances, assembled in varying proportions — earth, air, fire, and water, which possessed four properties — hot, dry, wet, and cold.

The idea of atoms did not surface again until the 17th and 18th centuries.

---

**Law of Conservation of Mass**

- In 1661, Robert Boyle redefined an **element** as a substance that cannot be chemically broken down further.
- **Law of Conservation of Mass** — Mass is neither created nor destroyed in chemical reactions (i.e., the total mass of a system does not change during a reaction). (Antoine Lavoisier, 1743-1794)
**Law of Definite Proportions**

- **Law of Definite Proportions** — All samples of a pure chemical substance, regardless of their source or how they were prepared, have the same proportions by mass of their constituent elements. (Joseph Proust, 1754-1826)
  
  - Calcium carbonate, which is found in coral, seashells, marble, limestone, chalk, and San Angelo tap water, is always 40.04% by mass calcium, 12.00% carbon, and 47.96% oxygen. (We now know that this results from the fact that calcium carbonate is CaCO$_3$.)

**The Law of Multiple Proportions**

- **Law of Multiple Proportions** — Elements can combine in different ways to form different substances, whose mass ratios are small whole-number multiples of each other. (John Dalton, 1804)

<table>
<thead>
<tr>
<th>Compound</th>
<th>Sample Size</th>
<th>Mass of Sulfur</th>
<th>Mass of Oxygen</th>
</tr>
</thead>
<tbody>
<tr>
<td>Sulfur oxide I</td>
<td>2.00 g</td>
<td>1.00 g</td>
<td>1.00 g</td>
</tr>
<tr>
<td>Sulfur oxide II</td>
<td>2.50 g</td>
<td>1.00 g</td>
<td>1.50 g</td>
</tr>
</tbody>
</table>

\[
\frac{\text{mass of oxygen in sulfur oxide II per gram of sulfur}}{\text{mass of oxygen in sulfur oxide I per gram of sulfur}} = \frac{1.50 \text{ g}}{1.00 \text{ g}} = \frac{3}{2}
\]
Dalton’s Atomic Theory

• John Dalton (1766-1844) explained these observations in 1808 by proposing the atomic theory:
  – Each element consists of tiny indivisible (not quite) particles called atoms.
  – All atoms of the same element have the same mass (not quite), but atoms of different elements have different masses.
  – Atoms combine in simple, whole-number ratios to form compounds. A given compound always has the same relative numbers and types of atoms.
  – Atoms of one element cannot change into atoms of another element (not quite). In a chemical reaction, atoms change the way they are bound to other atoms, but the atoms themselves are unchanged.

Dalton’s Atomic Theory

• Dalton’s atomic hypothesis had an uphill struggle — many scientists didn’t like the idea of using small, invisible entities to explain phenomena.
• Most (but not all) chemists had accepted the existence of atoms by the early 20th century; however, many influential physicists did not accept the atomic theory until Einstein’s landmark paper on Brownian motion (1905).
• Dalton’s original formulation of atoms as miniature billiard balls was incomplete: it did not explain how atoms combined to form compounds, or anything about their interior structure. The theory was modified greatly once charged particles coming from inside the atom (radioactivity) were discovered in the late 19th century.
The Electron

• In 1897, J. J. Thomson (1856-1940) investigated cathode rays, produced by passing an electric current through two electrodes in a vacuum tube (a cathode ray tube, CRT).

• The beam was produced at the negative electrode (cathode), and was deflected by the negative pole of an applied electrical field, implying that the rays were composed of negatively charged particles, with a very low mass. These particles were named electrons.

Figure 2.3

• Thomson’s experiments showed that electrons were emitted by many different types of metals, so electrons must be present in all types of atoms.

• Although Thomson was unable to measure the mass of the electron directly, he was able to determine the charge-to-mass ratio, $e/m$, $-1.758820 \times 10^8$ C/g.

  – This meant that the electron was about 2000 times lighter than hydrogen, the lightest element, and atoms were thus not the smallest unit of matter.
Chapter 2: Atoms, Ions, and Molecules

The Mass of the Electron

- In 1909, Robert Millikan (1868-1953) measured the charge on the electron by observing the movement of tiny ionized droplets of oil passing between two electrically charged plates. Knowing the $e/m$ ratio from Thomson’s work, the mass of the electron could then be determined:

**Charge of an electron:**

$$ e = -1.60218 \times 10^{-19} \text{ C} $$

**Charge to mass ratio:**

$$ e/m = -1.758820 \times 10^8 \text{ C/g} $$

**Mass of an electron:**

$$ m_e = 9.1093897 \times 10^{-28} \text{ g} $$

**Okay, Where’s the Positive Charge?**

- If there is negatively particle inside an electrically neutral atom, there must also be a positive charge.

- The model for the atom that Thomson proposed was of a diffuse, positively charged lump of matter with electrons embedded in it like “raisins in a plum pudding” (a watermelon or a blueberry muffin might be a more familiar analogy).
Radioactivity

• In the late 19th century, it was discovered that certain elements produce high-energy radiation.
  – In 1896, Henri Becquerel [Nobel Prize, 1903 (Phys.)] found that uranium produces an image on a photographic plate in the absence of light.
  – Marie Curie [Nobel Prize, 1903 (Phys.) and 1911 (Chem.)] and Pierre Curie [Nobel Prize, 1903 (Phys.)] discovered radioactivity in thorium, and isolated previously unknown elements (radium, polonium) that were even more radioactive.

• There are three major types of radiation:
  – alpha (α) particles — consists of two protons and two neutrons (a helium nucleus), having a +2 charge and a mass 7300 times that of an electron.
  – beta (β) particles — a high-speed electron emitted from the nucleus of an atom (when a neutron turns into a proton).
  – gamma (γ) rays — high-energy electromagnetic radiation.

The Discovery of the Nucleus

• In 1910, Ernest Rutherford [Nobel Prize, 1908, Chemistry] tested the “plum-pudding” model of the atom by firing a stream of alpha particles at a thin sheet of gold foil (about 2000 atoms thick).

• In the “plum-pudding” model, the mass of the atom is spread evenly through the volume of the atom. All of the alpha particles should plow right through the foil — but that’s not what happened . . .
The Discovery of the Nucleus

- ... instead, while most of the alpha-particles sailed through the gold foil, some were deflected at large angles, as if they had hit something massive, and some even bounced back toward the emitter.

It was quite the most incredible event that has ever happened to me in my life. It was almost as incredible as if you fired a 15-inch shell at a piece of tissue paper and it came back and hit you.

— Ernest Rutherford, in E. N. da C. Andrade, Rutherford and the Nature of the Atom (1964)

The Nuclear Atom Model

- Rutherford concluded that all of the positive charge and most of the mass (~99.9%) of the atom was concentrated in the center, called the nucleus. Most of the volume of the atom was empty space, through which the electrons were dispersed in some fashion.

- The positively charged particles within the nucleus are called protons; there must be one electron for each proton for an atom to be electrically neutral.

- This did not account for all of the mass of the atom, or the existence of isotopes (more later); the inventory of subatomic particles was “completed” (for the moment) by James Chadwick in 1932 [Nobel Prize, 1935], who discovered the neutron, an uncharged particle with about the same mass as the proton, which also resides in the nucleus.
The Modern View of Atomic Structure

The Atomic Theory Today

- An **atom** is an electrically neutral, spherical entity composed of a positively charged central **nucleus** surrounded by negatively charged **electrons**.
- The nucleus contains the **protons**, which have positive charges, and **neutrons**, which are neutral.
  - Neutrons are very slightly heavier than protons; protons are 1836 times heavier than electrons.
- The nucleus contains about 99.97% of the atom’s mass, but occupies 1 ten-trillionth of the its volume.
- The **electrons** (e⁻), which have negative charges, surround the nucleus, and account for most of the atomic volume.
  - *The number of electrons equals the number of protons in the nucleus of a neutral atom.*
The Atom and the Subatomic Particles

<table>
<thead>
<tr>
<th>Particle</th>
<th>Mass in kilograms (kg)</th>
<th>Mass in atomic mass units (amu)</th>
<th>Charge in Coulombs (C)</th>
<th>Relative charge</th>
</tr>
</thead>
<tbody>
<tr>
<td>Electron ((^0_1e))</td>
<td>9.10938\times10^{-31} kg</td>
<td>5.4858\times10^{-4} amu</td>
<td>-1.60218\times10^{-19} C</td>
<td>-1</td>
</tr>
<tr>
<td>Proton ((^1_1p))</td>
<td>1.67262\times10^{-27} kg</td>
<td>1.00728 amu</td>
<td>+1.60218\times10^{-19} C</td>
<td>+1</td>
</tr>
<tr>
<td>Neutron ((^0_0n))</td>
<td>1.67493\times10^{-27} kg</td>
<td>1.00867 amu</td>
<td>0</td>
<td>0</td>
</tr>
</tbody>
</table>

Figure 2.7

Atomic Number, Electrons

- What makes elements different from each another is the number of protons in their atoms, called the atomic number \((Z)\). All atoms of the same element contain the same number of protons.
  - The number of protons determines the number of electrons in a neutral atom.
  - Since most of the volume of the atom is taken up by the electrons, when two atoms interact with each other, it is the outermost (valence) electrons that are making contact with each other.
  - The number and arrangement of the electrons in an atom determines its chemical properties. Thus, the chemistry of an atom arises from its electrons.
Mass Number, and Isotopes

- The **mass number** (A) is the sum of the number of protons (Z) and neutrons (N) in the nucleus of an atom: \( A = Z + N \).
- **Isotopes** of an element have the same # of protons, but different #'s of neutrons.
  - Isotopes of an element have nearly identical chemical behavior.
  - A **nuclide** is the nucleus of an element with a particular combination of protons and neutrons.
  - A particular nuclide can be indicated by writing the name or symbol of the atom followed by a dash and the mass number (e.g., hydrogen-1).

Atomic Symbols

- The **atomic symbol** specifies information about the nuclear mass, atomic number, and charge on a particular element. Every element has a one- or two-letter symbol based on its English or Latin name.

\[
\text{atomic symbol} \downarrow \\
\begin{array}{c}
\text{mass number} \\
(\text{protons} + \text{neutrons})
\end{array} \quad \\
\begin{array}{c}
\text{atomic number} \\
(\text{protons})
\end{array} \quad \\
\begin{array}{c}
\text{charge}
\end{array} \\
\begin{array}{c}
\text{number of units in a molecule}
\end{array} \\
\begin{array}{c}
\text{Always capitalized!}
\end{array} \quad \\
\begin{array}{c}
\text{Never capitalized!}
\end{array}
\]

\[
\text{CO} \neq \text{Co} !!!!
\]
Chapter 2: Atoms, Ions, and Molecules

**Mass Number, and Isotopes**

Hydrogen-1
1 proton, 0 neutrons  
\[ Z = 1 \]
\[ A = 1 \]

Hydrogen-2 (deuterium)
1 proton, 1 neutron  
\[ Z = 1 \]
\[ A = 2 \]

Hydrogen-3 (tritium)
1 proton, 2 neutrons  
\[ Z = 1 \]
\[ A = 3 \]

\[ ^1_1H \]

**Ions**

- Neutral atoms have the same number of electrons as protons. In many chemical reactions, atoms gain or lose electrons to form charged particles called ions.
- For example, sodium loses one electron, resulting in a particle with 11 protons and 10 electrons, having a +1 charge:
  \[ \text{Na} \rightarrow \text{Na}^+ + e^- \]
- Positively charged ions are called cations.
- Fluorine gains one electron, resulting in a particle with 9 protons and 10 electrons, having a -1 charge:
  \[ \text{F} + e^- \rightarrow \text{F}^- \]
- Negatively charged ions are called anions.
Chapter 2: Atoms, Ions, and Molecules

**Examples: Writing Element Symbols**

1. Carbon-12 has how many protons? How many neutrons? How many electrons?
2. What would be the symbol for an element which has 14 protons and 15 neutrons?
3. What would be the symbol for an element which has 24 protons and 28 neutrons?
4. What would be the symbol for an element with 7 protons, 7 neutrons, and 10 electrons?
5. What would be the symbol for an element with 12 protons, 12 neutrons, and 10 electrons?
6. How many protons, neutrons, and electrons are there in $^{238}_{92}\text{U}$?
7. How many protons, neutrons, and electrons are in the $^{56}_{26}\text{Fe}^{3+}$ ion?

**Atomic Mass Units**

- The **average atomic mass** of an element is usually written underneath the element symbol on the periodic table.
- The masses of atoms are measured relative to the carbon-12 isotope, which is defined as weighing *exactly* 12 atomic mass units (amu, or dalton, Da).
  - 1 amu = 1 dalton = $1.660539 \times 10^{-24}$ g.
  - Protons and neutrons each weigh about 1 amu.
  - (Using carbon-12 as a reference allows the masses of other elements to be fairly close to whole numbers.)
- The **isotopic mass** of a particular isotope is mass of one atom of that isotope measured in amu’s. (Hydrogen-1 = 1.007825035 amu, hydrogen-2 = 2.014101779 amu.)
Atomic Masses

- When considering a sample of an element found in nature, we must take into account that the sample probably contains a number of different isotopes of the element.
  - For instance, hydrogen is mostly $^1\text{H}$ (99.985%), but there is also a small percentage of $^2\text{H}$ (deuterium, 0.015%).

- The atomic mass (or atomic weight) of an element is the average of the masses of the naturally-occurring isotopes of that element, weighted according to the isotopes’ abundance.
  - This number is obtained by adding up the weights of all the naturally occurring isotopes multiplied by their relative abundances.

For hydrogen,

\[
0.99985 \times 1.007825035 \text{ amu} = 1.0077 \text{ amu}
\]

\[
0.00015 \times 2.014101779 \text{ amu} = 0.00030 \text{ amu}
\]

\[
1.0080 \text{ amu}
\]

- This data can be obtained from a device called a mass spectrometer.
8. Use the following data to calculate the atomic mass of neon.

<table>
<thead>
<tr>
<th>Isotope</th>
<th>Mass</th>
<th>Abundance</th>
</tr>
</thead>
<tbody>
<tr>
<td>neon-20</td>
<td>19.992 amu</td>
<td>90.48%</td>
</tr>
<tr>
<td>neon-21</td>
<td>20.994 amu</td>
<td>0.27%</td>
</tr>
<tr>
<td>neon-22</td>
<td>21.991 amu</td>
<td>9.25%</td>
</tr>
</tbody>
</table>

Solution:

\[
0.9048 \times 19.992 \text{ amu} = 18.09 \text{ amu} \\
0.0027 \times 20.994 \text{ amu} = 0.057 \text{ amu} \\
0.0925 \times 21.991 \text{ amu} = 2.03 \text{ amu} \\
\text{20.177 amu} \\
\text{20.18 amu}
\]
The Elements

• All of the substances in the world are made of one or more of 118 elements, 92 of which occur naturally.

• An **element** is a substance which cannot be chemically broken down into simpler substances. Elements are defined by the number of protons in the nucleus.

• The elements are all assigned one or two letter symbols. The first letter is *always* capitalized, the second is *never* capitalized.

• The names, symbols, and other information about the 118 elements are organized into a chart called the **periodic table of the elements**.
### Names and Symbols of Some Common Elements

<table>
<thead>
<tr>
<th>Element</th>
<th>Symbol</th>
<th>Element</th>
<th>Symbol</th>
<th>Element</th>
<th>Symbol</th>
</tr>
</thead>
<tbody>
<tr>
<td>Aluminum</td>
<td>Al</td>
<td>Iodine</td>
<td>I</td>
<td>Antimony</td>
<td>Sb</td>
</tr>
<tr>
<td>Argon</td>
<td>Ar</td>
<td>Lithium</td>
<td>Li</td>
<td>Copper</td>
<td>Cu</td>
</tr>
<tr>
<td>Barium</td>
<td>Ba</td>
<td>Magnesium</td>
<td>Mg</td>
<td>Iron</td>
<td>Fe</td>
</tr>
<tr>
<td>Boron</td>
<td>B</td>
<td>Manganese</td>
<td>Mn</td>
<td>Gold</td>
<td>Au</td>
</tr>
<tr>
<td>Bromine</td>
<td>Br</td>
<td>Neon</td>
<td>Ne</td>
<td>Lead</td>
<td>Pb</td>
</tr>
<tr>
<td>Calcium</td>
<td>Ca</td>
<td>Nickel</td>
<td>Ni</td>
<td>Mercury</td>
<td>Hg</td>
</tr>
<tr>
<td>Carbon</td>
<td>C</td>
<td>Nitrogen</td>
<td>N</td>
<td>Silver</td>
<td>Ag</td>
</tr>
<tr>
<td>Chlorine</td>
<td>Cl</td>
<td>Oxygen</td>
<td>O</td>
<td>Sodium</td>
<td>Na</td>
</tr>
<tr>
<td>Chromium</td>
<td>Cr</td>
<td>Phosphorus</td>
<td>P</td>
<td>Potassium</td>
<td>K</td>
</tr>
<tr>
<td>Cobalt</td>
<td>Co</td>
<td>Silicon</td>
<td>Si</td>
<td>Tin</td>
<td>Sn</td>
</tr>
<tr>
<td>Fluorine</td>
<td>F</td>
<td>Sulfur</td>
<td>S</td>
<td>Tungsten</td>
<td>W</td>
</tr>
<tr>
<td>Helium</td>
<td>He</td>
<td>Titanium</td>
<td>Ti</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Hydrogen</td>
<td>H</td>
<td>Zinc</td>
<td>Zn</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

---

### Relative Abundances of the Elements

- **Element Abundances in the Universe**
- **Element Abundances in the Earth’s Crust**
- **Element Abundances in the Human Body**
Periodic Properties

- It has long been known that many of the elements have similar chemical properties.
  - Lithium, sodium, and potassium all perform the same reaction with water,
    \[2M(s) + 2\text{HOH}(l) \rightarrow 2\text{MOH}(aq) + \text{H}_2(g)\]
    the only difference being the masses of the metals themselves and the vigor and speed of the reaction.

The Invention of the Periodic Table

- In 1869 Dmitri Mendeleev published a table in which the elements that were known at the time were arranged by increasing atomic mass, and grouped into columns according to their chemical properties. The properties of the elements varied (more or less) in a periodic way in this arrangement.
Mendeleev’s Periodic Table

- Mendeleev noticed that when he grouped the elements by their properties, there were some “holes” which he guessed corresponded to as-yet-unknown elements.

- Mendeleev predicted some of the properties for two of these, eka-aluminum (atomic weight = 68), and eka-silicon (atomic weight = 72), which corresponded well to gallium (Ga, discovered in 1875) and germanium (Ge, 1886).

Mendeleev’s Predictions for Eka-Silicon vs. Observations for Germanium

<table>
<thead>
<tr>
<th>Property</th>
<th>Prediction for Eka-Silicon</th>
<th>Actual Properties of Germanium</th>
</tr>
</thead>
<tbody>
<tr>
<td>Atomic weight</td>
<td>72</td>
<td>72.3</td>
</tr>
<tr>
<td>Density</td>
<td>5.5 g cm⁻³</td>
<td>5.47 g cm⁻³</td>
</tr>
<tr>
<td>Specific heat</td>
<td>0.31 J g⁻¹ °C⁻¹</td>
<td>0.32 J g⁻¹ °C⁻¹</td>
</tr>
<tr>
<td>Melting point</td>
<td>high</td>
<td>960°C</td>
</tr>
<tr>
<td>Oxide formula</td>
<td>RO₂</td>
<td>GeO₂</td>
</tr>
<tr>
<td>Oxide density</td>
<td>4.7 g cm⁻³</td>
<td>4.70 g cm⁻³</td>
</tr>
<tr>
<td>Chloride formula</td>
<td>RCl₄</td>
<td>GeCl₄</td>
</tr>
<tr>
<td>Chloride boiling point</td>
<td>100°C</td>
<td>96°C</td>
</tr>
</tbody>
</table>
**The Periodic Table by Atomic Number**

- To make the properties of the elements “line up” properly, it was sometimes necessary to exchange the order of the elements.
  - For instance, potassium (39.0983 g/mol) is slightly lighter than argon (39.948 g/mol), so by increasing atomic weight, potassium should be in Group 8A, and argon in Group 1A, but that clearly doesn’t fit their observed properties.
- After the discovery of the nucleus and the proton, and with the development of X-ray spectroscopy, it was discovered that the periodic table could be written in order of **increasing atomic number**, with no need to “play around” with the order of the elements. It was also possible to count protons, and see exactly how many “missing” elements there were.

**Elements and the Periodic Table**

- The modern periodic table of the elements places the elements on a grid with 7 horizontal rows, called **periods**, and 18 vertical columns, called **groups**.
  - The elements are listed in order of increasing **atomic number**.
  - Two rows that are a part of periods 6 and 7 are shown beneath the table.
  - When they are organized in this way, there is a periodic pattern to the properties of the elements: **elements in the same group have similar chemical properties**.
  - The arrangement of the elements on the periodic table is a reflection of the interior structure of the atom (more later).
Chapter 2: Atoms, Ions, and Molecules

Group Numbers

- The **group number** can be written in a couple of different ways:
  - **1-18** is the IUPAC-recommended numbering system. This is more unambiguous, but less useful.
  - **1A-8A** for tall columns, **1B-8B** for short columns is the more commonly used numbering system.
- In the 1A-8A columns, the column numbers represent the number of *valence (outermost) electrons* for the main-group elements.
- The number of valence electrons are what primarily determines an atom’s chemistry.

![The Periodic Table of the Elements](http://www.angelo.edu/faculty/kboudrea/periodic.htm)

---

**Figure 2.12**

[Other Tables](http://www.angelo.edu/faculty/kboudrea/periodic.htm)
Parts of the Periodic Table — Main Groups

- **Main groups** (aka *representative elements*) — Groups 1A-8A (the tall columns); these elements have properties that are relatively predictable based on their positions on the table.

  - **Group 1A**, the *alkali metals* — lustrous, soft metals that react rapidly with water to make basic (alkaline) products. These elements are highly reactive, and are found in nature in compounds, and not in their elemental forms.

    - (Even though it is at the top of Group 1A, H is not considered an alkali metal.)
- **Group 2A**, the *alkaline earth metals* — lustrous, silvery, reactive metals. They are less reactive than the alkali metals, but are still too reactive to be found in the elemental form.

- **Group 7A**, the *halogens* — colorful, corrosive nonmetals; found in nature only in compounds.

- **Group 8A**, the *noble (inert) gases* — monatomic gases that are chemically stable and very unreactive.

---

### Parts of the Periodic Table — Transition Metals

- **Transition metal groups** — Groups 1B-8B (the shorter columns) — these metals exhibit a very wide range of properties, colors, reactivities, etc.

- **Inner transition metal groups** — these elements belong between groups 3B and 4B, but are usually shown tucked underneath the main table:
  - **Lanthanides** — elements 58-71 (following the element lanthanum, La). Most of these are not commonly known, although some have industrial and research applications. (Also called the “rare earth elements.”)
  - **Actinides** — elements 90-103 (following the element actinium, Ac). Most of these elements are either highly radioactive, or are synthesized in particle accelerators.
Chapter 2: Atoms, Ions, and Molecules

**Metals and Nonmetals**

- A jagged line on the periodic table separates the **metals** (left) from the **nonmetals** (right):

  - **Metals** are shiny, lustrous solids at room temperature (except for Hg, which is a liquid)
    - good conductors of electricity and heat.
    - malleable (can be hammered into thin sheets).
    - ductile (can be drawn into wire).
    - tend to lose electrons (oxidation) to form cations.

  ![Periodic Table Diagram](image)

  ![Metal and Nonmetal Dividing Line](image)

- **Nonmetals** are usually found in compounds, but some pure elemental forms are well-known: \( \text{N}_2, \text{O}_2, \text{C} \) (graphite and diamond), \( \text{Cl}_2 \), etc.
  - no metallic luster; not malleable or ductile.
  - poor conductors of electricity and heat.
  - tend to gain electrons (reduction) to form anions.

- Along the dividing line are the **semimetals** (or **metalloids**), which have properties intermediate between metals and nonmetals.
  - most of their physical properties resemble nonmetals.
  - several of the metalloids are **semiconductors**, which conduct electricity under special circumstances (Si, Ge).
Chapter 2: Atoms, Ions, and Molecules

**Metals and Nonmetals**

- Most things that we encounter are not elements, but **compounds**, composed of two or more elements.
  - **Binary compounds** are composed of two elements (H₂O, CH₄, NH₃, NaCl, CaCl₂, etc.)
  - **Diatomic compounds** are composed of two atoms, which may or may not be the same (Cl₂, CO, etc.)

- There are two major types of chemical compounds:
  - **molecular compounds** — nonmetal + nonmetal
    - held together by covalent bonds that result from the *sharing* of pairs of electrons.
  - **ionic compounds** — metal + nonmetal
    - held together by **ionic bonds**, which result from the *transfer* of electrons from the metal to the nonmetal, producing **ions**.
Ions and the Periodic Table

- Main group metals tend to lose electrons to form cations that have the same number of electrons as the preceding noble gas. **The charge on the typical cation is the same as the group number.**
  - Group 1A: +1  Group 3A: +3
  - Group 2A: +2
- Main group nonmetals tend to gain electrons to form anions that have with the same number of electrons as the nearest noble gas. **The charge on the typical anion is the group number minus eight.**
  - Group 5A: -3  Group 7A: -1
  - Group 6A: -2
- This is known as the **octet rule** — more later
Formulas of Ionic Compounds

• The smallest unit of an ionic compound is the **formula unit**, the smallest *electrically neutral* collection of ions (NaCl, CaCl₂, Na₂S, Al₂O₃, etc.)

• **Monatomic ions** are cations or anions derived from a single atom, such as Cl⁻, O²⁻, Na⁺, and Mg²⁺.

• **Polyatomic ions** are combinations of atoms that possess an overall charge, such as CO₃²⁻, SO₄²⁻, NO₃⁻, CN⁻, NH₄⁺, C₂H₃O₂⁻, etc.

---

Examples: Writing Ionic Formulas

1. Write the formula for the ionic compound formed between the following pairs of elements and provide a name for the compound.

   a. Al and F
   b. Na and S
   c. Ba and S
   d. Mg and P
   e. Ca and Cl
   f. Na and P
## Naming Chemical Compounds

### Main-Group Metals

- Group 1A, 2A, and 3A metals tend to form *cations* by losing all of their outermost (valence) electrons.
- The charge on the cation is the same as the group number.
- The cation is given the same name as the neutral metal atom, with the word “ion” added to the end.

<table>
<thead>
<tr>
<th>Group</th>
<th>Ion</th>
<th>Ion name</th>
<th>Group</th>
<th>Ion</th>
<th>Ion name</th>
</tr>
</thead>
<tbody>
<tr>
<td>1A</td>
<td>H⁺</td>
<td>hydrogen ion</td>
<td>2A</td>
<td>Mg²⁺</td>
<td>magnesium ion</td>
</tr>
<tr>
<td></td>
<td>Li⁺</td>
<td>lithium ion</td>
<td></td>
<td>Ca²⁺</td>
<td>calcium ion</td>
</tr>
<tr>
<td></td>
<td>Na⁺</td>
<td>sodium ion</td>
<td></td>
<td>Sr²⁺</td>
<td>strontium ion</td>
</tr>
<tr>
<td></td>
<td>K⁺</td>
<td>potassium ion</td>
<td></td>
<td>Ba²⁺</td>
<td>barium ion</td>
</tr>
<tr>
<td></td>
<td>Cs⁺</td>
<td>cesium ion</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>3A</td>
<td>Al³⁺</td>
<td>aluminum ion</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
**Transition and Post-Transition Metals**

- Many transition and post-transition metals can form cations with more than one possible charge.
  - The common charges of the transition metals must be memorized.
  - The charges of the Group 4A and 5A metal cations are either the group number, or the group number minus two.

- **Common or trivial names:** -ic endings go with the higher charge, -ous endings go with the lower charge.
  - Often, the name used is the Latin name of the element (e.g., iron = ferrum)
  - \( \text{Fe}^{2+} \) ferrous ion, \( \text{Fe}^{3+} \) ferric ion
  - \( \text{Cu}^+ \) cuprous ion, \( \text{Cu}^{2+} \) cupric ion

**Transition and Post-Transition Metals**

- **Systematic names (Stock system):** name the metal, followed by the charge in parentheses (written in Roman numerals).
  - \( \text{Fe}^{2+} \) iron(II) ion, \( \text{Fe}^{3+} \) iron(III) ion
  - \( \text{Cu}^+ \) copper(I) ion, \( \text{Cu}^{2+} \) copper(II) ion
  - Roman numerals should be used on all transition metals and post-transition metals except for \( \text{Ag}^+ \), \( \text{Cd}^{2+} \), and \( \text{Zn}^{2+} \).
## Transition and Post-Transition Metals

<table>
<thead>
<tr>
<th>Ion</th>
<th>Systematic name</th>
<th>Common name</th>
</tr>
</thead>
<tbody>
<tr>
<td>Cr$^{2+}$</td>
<td>chromium(II) ion</td>
<td>chromous ion</td>
</tr>
<tr>
<td>Cr$^{3+}$</td>
<td>chromium(III) ion</td>
<td>chromic ion</td>
</tr>
<tr>
<td>Mn$^{2+}$</td>
<td>manganese(II) ion</td>
<td>manganous ion</td>
</tr>
<tr>
<td>Mn$^{3+}$</td>
<td>manganese(III) ion</td>
<td>manganic ion</td>
</tr>
<tr>
<td>Fe$^{2+}$</td>
<td>iron(II) ion</td>
<td>ferrous ion</td>
</tr>
<tr>
<td>Fe$^{3+}$</td>
<td>iron(III) ion</td>
<td>ferric ion</td>
</tr>
<tr>
<td>Co$^{2+}$</td>
<td>cobalt(II) ion</td>
<td>cobaltous ion</td>
</tr>
<tr>
<td>Co$^{3+}$</td>
<td>cobalt(III) ion</td>
<td>cobaltic ion</td>
</tr>
<tr>
<td>Ni$^{2+}$</td>
<td>nickel(II) ion</td>
<td></td>
</tr>
<tr>
<td>Cu$^{+}$</td>
<td>copper(I) ion</td>
<td>cuprous ion</td>
</tr>
<tr>
<td>Cu$^{2+}$</td>
<td>copper(II) ion</td>
<td>cupric ion</td>
</tr>
<tr>
<td>Zn$^{2+}$</td>
<td>zinc ion</td>
<td></td>
</tr>
<tr>
<td>Ag$^{+}$</td>
<td>silver ion</td>
<td></td>
</tr>
<tr>
<td>Cd$^{2+}$</td>
<td>cadmium ion</td>
<td></td>
</tr>
<tr>
<td>Au$^{3+}$</td>
<td>gold(III) ion</td>
<td></td>
</tr>
<tr>
<td>Hg$^{2+}$</td>
<td>mercury(I) ion</td>
<td>mercurous ion</td>
</tr>
<tr>
<td>Hg$^{2+}$</td>
<td>mercury(II) ion</td>
<td>mercuric ion</td>
</tr>
<tr>
<td>Sn$^{2+}$</td>
<td>tin(II) ion</td>
<td>stannous ion</td>
</tr>
<tr>
<td>Sn$^{4+}$</td>
<td>tin(IV) ion</td>
<td>stannic ion</td>
</tr>
<tr>
<td>Pb$^{2+}$</td>
<td>lead(II) ion</td>
<td>plumbous ion</td>
</tr>
<tr>
<td>Pb$^{4+}$</td>
<td>lead(IV) ion</td>
<td>plumbic ion</td>
</tr>
<tr>
<td>Bi$^{3+}$</td>
<td>bismuth(III) ion</td>
<td></td>
</tr>
<tr>
<td>Bi$^{5+}$</td>
<td>bismuth(V) ion</td>
<td></td>
</tr>
</tbody>
</table>
Main-Group Nonmetals

- Group 4A - 7A nonmetals form *anions* by gaining enough electrons to fill their valence shell (eight electrons). The charge on the anion is the group number minus eight.

- The anion is named by taking the element stem and adding the ending *-ide*.

<table>
<thead>
<tr>
<th>Group</th>
<th>Ion</th>
<th>Ion name</th>
<th>Group</th>
<th>Ion</th>
<th>Ion name</th>
</tr>
</thead>
<tbody>
<tr>
<td>4A</td>
<td>C⁴⁻</td>
<td>carbide ion</td>
<td>6A</td>
<td>Se²⁻</td>
<td>selenide ion</td>
</tr>
<tr>
<td></td>
<td>Si⁴⁻</td>
<td>silicide ion</td>
<td></td>
<td>Te²⁻</td>
<td>telluride ion</td>
</tr>
<tr>
<td>5A</td>
<td>N³⁻</td>
<td>nitride ion</td>
<td>7A</td>
<td>F⁻</td>
<td>fluoride ion</td>
</tr>
<tr>
<td></td>
<td>P³⁻</td>
<td>phosphide ion</td>
<td></td>
<td>Cl⁻</td>
<td>chloride ion</td>
</tr>
<tr>
<td></td>
<td>As³⁻</td>
<td>arsenide ion</td>
<td></td>
<td>Br⁻</td>
<td>bromide ion</td>
</tr>
<tr>
<td>6A</td>
<td>O²⁻</td>
<td>oxide ion</td>
<td>1A</td>
<td>H⁻</td>
<td>hydride ion</td>
</tr>
<tr>
<td></td>
<td>S²⁻</td>
<td>sulfide ion</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Common Cations and Anions

<table>
<thead>
<tr>
<th>Elements To Memorize</th>
<th>IIA</th>
<th>IIB</th>
<th>IIIA</th>
<th>IIIB</th>
<th>IVB</th>
<th>VB</th>
<th>VIB</th>
<th>VIIB</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydrogen</td>
<td>Li</td>
<td>Be</td>
<td>B</td>
<td>C</td>
<td>N</td>
<td>O</td>
<td>F</td>
<td>Ne</td>
</tr>
<tr>
<td>Helium</td>
<td>Na</td>
<td>Mg</td>
<td>Al</td>
<td>Si</td>
<td>P</td>
<td>S</td>
<td>Cl</td>
<td>Ar</td>
</tr>
<tr>
<td>Sodium</td>
<td>K</td>
<td>Ca</td>
<td>Cr</td>
<td>Mn</td>
<td>Fe</td>
<td>Co</td>
<td>Ni</td>
<td>Cu</td>
</tr>
<tr>
<td>Calcium</td>
<td>Rb</td>
<td>Sr</td>
<td>Ag</td>
<td>Cd</td>
<td>Sn</td>
<td>Sb</td>
<td>Te</td>
<td>I</td>
</tr>
<tr>
<td>Rubidium</td>
<td>Cs</td>
<td>Ba</td>
<td>Au</td>
<td>Hg</td>
<td>Tl</td>
<td>Pb</td>
<td>Bi</td>
<td>Rn</td>
</tr>
<tr>
<td>Barium</td>
<td>Ra</td>
<td>Ac</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Lanthanides</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Actinides</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Lanthanides</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Actinides</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

61 62
### Polyatomic Ions

- **Polyatomic ions** are ions composed of groups of covalently bonded atoms which have an overall charge.

<table>
<thead>
<tr>
<th>Ion</th>
<th>Name</th>
<th>Formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>NH$_4^+$</td>
<td>ammonium</td>
<td>OCN$^-$ cyanate</td>
</tr>
<tr>
<td>H$_3$O$^+$</td>
<td>hydronium</td>
<td>MnO$_4^-$ permanganate</td>
</tr>
<tr>
<td>OH$^-$</td>
<td>hydroxide</td>
<td>C$_2$H$_3$O$_2^-$ acetate (OAc$^-$, CH$_3$CO$_2^-$)</td>
</tr>
<tr>
<td>CN$^-$</td>
<td>cyanide</td>
<td>CO$_3^{2-}$ carbonate</td>
</tr>
<tr>
<td>O$_2^{2-}$</td>
<td>peroxide</td>
<td>HCO$_3^-$ hydrogen carbonate, bicarbonate</td>
</tr>
<tr>
<td>N$_3^-$</td>
<td>azide</td>
<td>SO$_4^{2-}$ sulfate</td>
</tr>
<tr>
<td>NO$_3^-$</td>
<td>nitrate</td>
<td>SO$_3^{2-}$ sulfite</td>
</tr>
<tr>
<td>NO$_2^-$</td>
<td>nitrite</td>
<td>S$_2$O$_3^{2-}$ thiosulfate</td>
</tr>
<tr>
<td>ClO$_3^-$</td>
<td>chlorate</td>
<td>C$_2$O$_4^{2-}$ oxalate</td>
</tr>
<tr>
<td>ClO$_2^-$</td>
<td>chlorite</td>
<td>CrO$_4^{2-}$ chromate</td>
</tr>
<tr>
<td>ClO$^-$</td>
<td>hypochlorite</td>
<td>Cr$_2$O$_7^{2-}$ dichromate</td>
</tr>
<tr>
<td>ClO$_4^-$</td>
<td>perchlorate</td>
<td>PO$_4^{3-}$ phosphate</td>
</tr>
</tbody>
</table>

### Polyatomic Ions — Regularities in Names

- There are some regularities in the names of these polyatomic ions:
  - **Thio-** implies replacing an oxygen with a sulfur:
    - SO$_4^{2-}$ sulfate
    - OCN$^-$ cyanate
    - S$_2$O$_3^{2-}$ thiosulfate
    - SCN$^-$ thiocyanate

- Replacing the first element with another element from the same group gives a polyatomic ion with the same charge, and a similar name:

<table>
<thead>
<tr>
<th>Group 7A</th>
<th>Group 6A</th>
<th>Group 5A</th>
<th>Group 4A</th>
</tr>
</thead>
<tbody>
<tr>
<td>ClO$_3^-$</td>
<td>chlorate</td>
<td>SO$_4^{2-}$</td>
<td>sulfate</td>
</tr>
<tr>
<td>BrO$_3^-$</td>
<td>bromate</td>
<td>SeO$_4^{2-}$</td>
<td>selenate</td>
</tr>
<tr>
<td>IO$_3^-$</td>
<td>iodate</td>
<td>TeO$_4^{2-}$</td>
<td>tellurate</td>
</tr>
</tbody>
</table>
Polyatomic Ions — Oxoanions

- Some nonmetals form a series of oxoanions having different numbers of oxygens (all with the same charge). The general rule for such series is shown below. (Note that in some cases, the -ate form has three oxygens, and in some cases four oxygens. These forms must be memorized.)

\[
\begin{align*}
XO_n^- & \quad \text{stem + ate} \quad \text{ClO}_3^- \quad \text{chlorate} \\
XO_{n-1}^- & \quad \text{stem + ite} \quad \text{ClO}_2^- \quad \text{chlorite} \\
XO_{n-2}^- & \quad \text{hypo + stem + ite} \quad \text{ClO}^- \quad \text{hypochlorite} \\
XO_{n+1}^- & \quad \text{per + stem + ate} \quad \text{ClO}_4^- \quad \text{perchlorate} \\
X_y^- & \quad \text{stem + ide} \quad \text{Cl}^- \quad \text{chloride} \\
\end{align*}
\]

…the monatomic ion

Polyatomic Ions — Ions Containing Hydrogens

- Acid salts are ionic compounds that still contain an acidic hydrogen, such as NaHSO₄. In naming these salts, specify the number of acidic hydrogens still in the salt.

- The prefix bi- implies an acidic hydrogen.

\[
\begin{align*}
\text{CO}_3^{2-} & \quad \text{carbonate} \\
\text{HCO}_3^- & \quad \text{hydrogen carbonate, bicarbonate} \\
\text{SO}_4^{2-} & \quad \text{sulfate} \\
\text{HSO}_4^- & \quad \text{hydrogen sulfate, bisulfate} \\
\text{PO}_4^{3-} & \quad \text{phosphate} \\
\text{HPO}_4^{2-} & \quad \text{monohydrogen phosphate} \\
\text{H}_2\text{PO}_4^- & \quad \text{dihydrogen phosphate} \\
\end{align*}
\]
Writing Formulas of Ionic Compounds

- The cation is written first, followed by the monatomic or polyatomic anion.
- The subscripts in the formula must produce an electrically neutral formula unit.
- The subscripts should be the smallest set of whole numbers possible.
- If there is only one of a polyatomic ion in the formula, do not place parentheses around it. If there is more than one of a polyatomic ion, put the ion in parentheses, and place the subscript after the parentheses.
  - Remember the Prime Directive for formulas:
    \[ \text{Ca(OH)}_2 \neq \text{CaOH}_2! \]

Nomenclature of Binary Ionic Compounds: Metal + Nonmetal

- A binary compound is a compound formed from two different elements. A diatomic compound (or diatomic molecule) contains two atoms, which may or may not be the same.
- Metals combine with nonmetals to form ionic compounds. Name the cation first (specify the charge, if necessary), then the nonmetal anion (element stem + -ide).
- Do NOT use counting prefixes! This information is implied in the name of the compound.

\[
\begin{array}{lll}
\text{name of metal cation} & \left(\text{charge of metal cation in Roman numerals in parenthesis (if necessary)}\right) & \text{element stem of nonmetal anion + -ide}
\end{array}
\]
**Nomenclature of Ionic Compounds:**

*Metal + Polyatomic Ion*

- Metals combine with polyatomic ions to give ionic compounds. Name the cation first (specify the charge, if necessary), then the polyatomic ion as listed in the previous table.
- Once again, do NOT use counting prefixes!

<table>
<thead>
<tr>
<th>name of metal cation</th>
<th>charge of metal cation in Roman numerals in parenthesis (if necessary)</th>
<th>name of polyatomic ion</th>
</tr>
</thead>
</table>

### Nomenclature of Ionic Compounds: Examples

<p>| | | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>Na</td>
<td>+</td>
<td>Cl</td>
</tr>
<tr>
<td>Na</td>
<td>+</td>
<td>S</td>
</tr>
<tr>
<td>Na</td>
<td>+</td>
<td>P</td>
</tr>
<tr>
<td>Ca</td>
<td>+</td>
<td>Cl</td>
</tr>
<tr>
<td>Ca</td>
<td>+</td>
<td>S</td>
</tr>
<tr>
<td>iron(II)</td>
<td>+</td>
<td>Cl</td>
</tr>
<tr>
<td>iron(III)</td>
<td>+</td>
<td>Cl</td>
</tr>
<tr>
<td>Na</td>
<td>+</td>
<td>sulfate</td>
</tr>
<tr>
<td>Ca</td>
<td>+</td>
<td>carbonate</td>
</tr>
<tr>
<td>Cr</td>
<td>+</td>
<td>nitrate</td>
</tr>
<tr>
<td>Ag</td>
<td>+</td>
<td>nitrite</td>
</tr>
</tbody>
</table>
Nomenclature of Ionic Compounds: Hydrates

- **Hydrates** are ionic compounds which also contain a specific number of water molecules associated with each formula unit. The water molecules are called *waters of hydration*.
- The formula for the ionic compound is followed by a raised dot and \#H\textsubscript{2}O — e.g., MgSO\textsubscript{4}·7H\textsubscript{2}O.
- They are named as ionic compounds, followed by a counting prefix and the word “hydrate”

\[
\begin{align*}
\text{MgSO}_4\cdot7\text{H}_2\text{O} & \quad \text{magnesium sulfate heptahydrate (Epsom salts)} \\
\text{CaSO}_4\cdot\frac{1}{2}\text{H}_2\text{O} & \quad \text{calcium sulfate hemihydrate} \\
\text{BaCl}_2\cdot6\text{H}_2\text{O} & \quad \text{barium chloride hexahydrate} \\
\text{CuSO}_4\cdot5\text{H}_2\text{O} & \quad \text{copper(II) sulfate pentahydrate}
\end{align*}
\]

Nomenclature of Binary Molecular Compounds: Nonmetal + Nonmetal

- Two nonmetals combine to form a *molecular* or *covalent compound* (i.e., one that is held together by covalent bonds, not ionic bonds).
- In many cases, two elements can combine in several ways to make completely different compounds (e.g., CO and CO\textsubscript{2}). It is necessary to specify how many of each element is present within the compound.
- In writing formulas, the more cation-like element (the one further to the left on the periodic table) is placed first, then the more anion-like element (the one further to the right on the periodic table).
- **Important exception:** halogens are written before oxygen. For two elements in the same group, the one with the higher period number is placed first.

\[
\begin{align*}
\end{align*}
\]
Nomenclature of Binary Molecular Compounds

- The first element in the formula is given the element name, and the second one is named by replacing the ending of the element name with \textit{-ide}.
- A numerical prefix is used in front of each element name to indicate how many of that element is present. (If there is only one of the first element in the formula, the \textit{mono-} prefix is dropped.)

\begin{center}
\begin{tabular}{|c|c|c|}
\hline
prefix & name of 1st element & stem of 2nd element + -ide \\
\hline
1 mono- & 4 tetra- & 7 hepta- \\
2 di- & 5 penta- & 8 octa- \\
3 tri- & 6 hexa- & 9 nona- \\
\hline
\end{tabular}
\end{center}

Nomenclature of Binary Molecular Compounds

- Some molecular compounds are known by common or trivial names:
  - NO \hspace{1em} nitrogen monoxide
  - NO$_2$ \hspace{1em} nitrogen dioxide
  - N$_2$O \hspace{1em} dinitrogen monoxide
  - N$_2$O$_3$ \hspace{1em} dinitrogen trioxide
  - N$_2$O$_4$ \hspace{1em} dinitrogen tetroxide
  - N$_2$O$_5$ \hspace{1em} dinitrogen pentoxide

- Some molecular compounds are known by common or trivial names:
  - H$_2$O \hspace{1em} water
  - NH$_3$ \hspace{1em} ammonia
1. Write the formula for the ionic compound formed between the following pairs of species and provide a name for the compound.

   a. Mg and phosphate ____________________
   b. Ammonium and nitrate ____________________
   c. Ammonium and sulfate ____________________
   d. Zn and Cl ____________________
   e. Mercury(I) and nitrite ____________________
   f. Mercury(II) and sulfite ____________________
   g. Chromium and S ____________________

2. Name the following compounds.

   a. Ca(NO$_3$)$_2$ ____________________
   b. BaCO$_3$ ____________________
   c. SO$_3$ ____________________
   d. SnCl$_4$ ____________________
   e. Fe$_2$(CO$_3$)$_3$ ____________________
   f. AlPO$_4$ ____________________
   g. N$_2$O ____________________
Examples: Formulas and Nomenclature

3. Name the following compounds.
   a. CrO  ___________________________
   b. Mn$_2$O$_3$ ______________________
   c. NO$_2$ __________________________
   d. NaNO$_2$ _________________________
   e. PBr$_3$ __________________________
   f. KHSO$_4$ _________________________
   g. LiH$_2$PO$_4$ ______________________

Examples: Formulas and Nomenclature

4. Write formulas for the following compounds.
   a. sodium nitrite  __________________ 
   b. lithium hydroxide  ______________
   c. barium chlorate  ________________
   d. potassium perchlorate  __________
   e. diphosphorus pentoxide  __________
   f. magnesium phosphate  ____________
   g. iron(II) carbonate  ______________
Examples: Formulas and Nomenclature

5. Write formulas for the following compounds.

   a. calcium bicarbonate  ________________
   b. manganese(III) carbonate  ________________
   c. potassium hypochlorite  ________________
   d. silver chromate  ________________
   e. nickel acetate  ________________
   f. barium peroxide  ________________
   g. titanium(IV) oxide  ________________

Nomenclature of Acids

- Acids are compounds in which the “cation” is H⁺. These are often given special “acid names” derived by omitting the word “hydrogen,” adding the word “acid” at the end, and changing the compound suffix as shown below:

<table>
<thead>
<tr>
<th>Compound name</th>
<th>Acid name</th>
</tr>
</thead>
<tbody>
<tr>
<td>oxoacids</td>
<td></td>
</tr>
<tr>
<td>stem + ate</td>
<td>stem + ic acid</td>
</tr>
<tr>
<td>stem + ite</td>
<td>stem + ous acid</td>
</tr>
<tr>
<td>binary acids</td>
<td></td>
</tr>
<tr>
<td>stem + ide</td>
<td>hydro + stem + ic acid</td>
</tr>
</tbody>
</table>

HClO₃ hydrogen chlorate  chloric acid
H₂SO₄ hydrogen sulfate  sulfuric acid
HClO₂ hydrogen chlorite  chlorous acid
HCl hydrogen chloride  hydrochloric acid
Chapter 2: Atoms, Ions, and Molecules

**Examples: Acid Nomenclature**

6. Write formulas or names for the following acids.

   a. HCl  ________________
   b. HClO₂  ________________
   c. H₂SO₃  ________________
   d. H₃PO₄  ________________
   e. hydrofluoric acid  ________________
   f. periodic acid  ________________
   g. chloric acid  ________________
   h. phosphorous acid  ________________

**Examples: Formulas and Nomenclature**

7. Which of the following formulas and/or names is written *incorrectly*?

   a. NaSO₄
   b. Na₂Cl
   c. MgNO₃
   d. magnesium dichloride
   e. iron(III) phosphate, Fe₃(PO₄)₂
   f. tin(IV) sulfate, Sn(SO₄)₂
   g. nitrogen chloride, NCl₃
   h. HClO₂, hypobromous acid