Phenomena: Some elements, such as “A” and “B”, are known to form several compounds. Data on the masses of A and B were collected on different sample sizes of the three compounds. What patterns do you notice about each of the compounds? What patterns do you notice between the three compounds?

<table>
<thead>
<tr>
<th>Substance x</th>
<th>Mass of A</th>
<th>Mass of B</th>
</tr>
</thead>
<tbody>
<tr>
<td>Sample 1</td>
<td>0.86 g</td>
<td>1.96 g</td>
</tr>
<tr>
<td>Sample 2</td>
<td>4.26 g</td>
<td>9.73 g</td>
</tr>
<tr>
<td>Sample 3</td>
<td>32.8 g</td>
<td>74.9 g</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Substance y</th>
<th>Mass of A</th>
<th>Mass of B</th>
</tr>
</thead>
<tbody>
<tr>
<td>Sample 1</td>
<td>2.84 g</td>
<td>3.24 g</td>
</tr>
<tr>
<td>Sample 2</td>
<td>5.10 g</td>
<td>5.83 g</td>
</tr>
<tr>
<td>Sample 3</td>
<td>0.86 g</td>
<td>0.98 g</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Substance z</th>
<th>Mass of A</th>
<th>Mass of B</th>
</tr>
</thead>
<tbody>
<tr>
<td>Sample 1</td>
<td>2.96 g</td>
<td>1.69 g</td>
</tr>
<tr>
<td>Sample 2</td>
<td>0.43 g</td>
<td>0.25 g</td>
</tr>
<tr>
<td>Sample 3</td>
<td>22.2 g</td>
<td>12.7 g</td>
</tr>
</tbody>
</table>
Big Idea: All matter is made of atoms. Atoms contain three fundamental particles. The number of protons in an atom defines the type of atom. Atoms join together to form larger compounds. These compounds belong to one of two classes: molecular or ionic.
How do we know there are atoms?

- **Law of Conservation of Mass**: $m_{\text{reactants}} = m_{\text{products}}$

- **Law of Definite Proportions**: When two or more elements combine to form a compound, their masses in that compound are in a fixed and definite ratio.

- **Law of Multiple Proportions**: When two elements form a series of compounds, the ratio of the masses of the second element that combined with 1 gram of the first element can always be reduced to small whole numbers.
Dalton’s Model (1808)

- Each element is made up of tiny particles called atoms.
- The atoms of a given element are identical; the atoms of different elements are different in some fundamental way.
- Chemical compounds are formed when atoms combine with each other. A given compound always has the same relative number and type of atoms.
- Chemical reactions involve reorganization of the atoms. The atoms themselves are not changed in a chemical reaction.
The e\textsuperscript{−} was discovered while investigating cathode rays which are the rays that are emitted when a high potential difference (high voltage) is applied between two electrodes in a glass tube. J.J. Thomson showed that cathode rays are streams of negatively charged particles. He found that regardless of the material he used, the same particles were found. He concluded that they are part of the makeup of all atoms we now call the electrons.

J. J. Thomson at the Cavendish laboratory in 1897

Chapter 2: Atoms, Molecules, and Ions
Robert Millikan carried out experiments that enabled the charge of the electron to be calculated. Since each drop contained more than one electron, he took the charge of the electron to be the smallest increment of charge between droplets. The modern value is \(-e\) with 
$$e = 1.602 \times 10^{-19} \text{ C}.$$
Ernest Rutherford knew that some elements emitted positively charged particles which he called α particles. Rutherford took a beam of α particles and shot it at a piece of Pt foil that was only a few atoms thick. The experiment showed that almost all the α particles passed through and were deflected only very slightly but about 1 in 20,000 were deflected by more than 90°. Therefore, the jelly model for atoms was wrong.
Atoms

Atoms are made up of three types of elementary particles:

<table>
<thead>
<tr>
<th>Particle</th>
<th>Symbol</th>
<th>Charge</th>
<th>Mass</th>
</tr>
</thead>
<tbody>
<tr>
<td>electron</td>
<td>e(^-)</td>
<td>-1.602×10(^{-19})</td>
<td>9.109×10(^{-31}) kg 0.00055 amu</td>
</tr>
<tr>
<td>proton</td>
<td>p</td>
<td>1.602×10(^{-19})</td>
<td>1.673×10(^{-27}) kg 1.0073 amu</td>
</tr>
<tr>
<td>neutron</td>
<td>n</td>
<td>0</td>
<td>1.675×10(^{-27}) kg 1.0087 amu</td>
</tr>
</tbody>
</table>

Parts of an Atom

- **Nucleus**: Contains protons and neutrons (positively charged, makes up most of the mass of the atom, \(\sim 10^{-13}\) cm in diameter)
- **Electron Cloud**: Contains electrons (negatively charged, 10\(^{-8}\) cm in diameter)
Atom Facts:
- # of protons = # of electrons (atoms are neutral)
- # of neutrons can vary

Isotope: Atoms with the same number of protons but different numbers of neutrons

What defines the type of atom?

**Atomic Number ($Z$):** Number of protons ($Z = p$)

*Note: This is the number on the periodic table*

**Mass Number ($A$):** Number of protons plus number of neutrons ($A = p + n$)
Atoms

Element-Mass Number
carbon-11

<table>
<thead>
<tr>
<th>p</th>
<th>e^-</th>
<th>n</th>
</tr>
</thead>
</table>

Short Hand Notation \( \text{Mass} \ # \ # \text{Element Symbol} \):

\( ^{20}_{10}\text{Ne} \text{ or } ^{20}\text{Ne} \)

<table>
<thead>
<tr>
<th>p</th>
<th>e^-</th>
<th>n</th>
</tr>
</thead>
</table>

Which of the following represent a pair of isotopes?

a) \(^{12}\text{C} \text{ and an atom with } p=12, \text{ and } n=12\)

b) An atom with atomic number=18 and mass number =40 and calcium-40

c) \(^{9}_{5}\text{B} \text{ and an atom with } p=4 \text{ and } n=5\)

d) \(^{26}\text{Mg} \text{ and an atom with } A=24 \text{ and } Z=12\)
Elements are arranged by increasing atomic number
Vertical columns are called groups
Groups 1, 2 and 13-18 are called the main group
Horizontal rows are called periods
The four rectangular regions of the periodic table are called s, p, d, and f blocks
Periodic Table

**Group 1 (Alkali Metals)**
- Properties: soft, lustrous, metallic, low melting temperatures, and highly reactive

**Group 2 (Alkaline Earth Metals)**
- Properties: similar to group 1 metals however they react less vigorously

**Groups 3-11 (Transition Metals)**
- Get their name because these elements are transitional in character between the vigorously reactive metals in the s block and the less reactive metals on the left of the p block

**Group 17 (Halogens)**
- Properties: all nonmetals, very reactive with alkali metals and alkaline earth metals

**Group 18 (Noble Gases)**
- Properties: colorless, odorless, gasses, which are inert due to their full octets
Metal: conducts electricity, has a luster, is malleable, and is ductile
Nonmetal: does not conduct electricity and is neither malleable nor ductile
Metalloid: Has the appearance and some properties of a metal but behaves chemically like a nonmetal
## Ions

### Neon

<table>
<thead>
<tr>
<th>Particle</th>
<th>Number</th>
</tr>
</thead>
<tbody>
<tr>
<td>Proton (+)</td>
<td>10</td>
</tr>
<tr>
<td>Neutron (0)</td>
<td>10</td>
</tr>
<tr>
<td>Electron (-)</td>
<td></td>
</tr>
<tr>
<td>Charge</td>
<td></td>
</tr>
</tbody>
</table>

### Fluorine

<table>
<thead>
<tr>
<th>Particle</th>
<th>Number</th>
</tr>
</thead>
<tbody>
<tr>
<td>Proton (+)</td>
<td>19</td>
</tr>
<tr>
<td>Neutron (0)</td>
<td>10</td>
</tr>
<tr>
<td>Electron (-)</td>
<td>9</td>
</tr>
<tr>
<td>Charge</td>
<td></td>
</tr>
</tbody>
</table>

### Fluorine Ion

<table>
<thead>
<tr>
<th>Particle</th>
<th>Number</th>
</tr>
</thead>
<tbody>
<tr>
<td>Proton (+)</td>
<td>19</td>
</tr>
<tr>
<td>Neutron (0)</td>
<td>9</td>
</tr>
<tr>
<td>Electron (-)</td>
<td>10</td>
</tr>
<tr>
<td>Charge</td>
<td></td>
</tr>
</tbody>
</table>

**Will F Gain or Lose e⁻?**
### Ion Properties

**Neon**
- **Proton (+):** 10
- **Neutron (0):** 10
- **Electron (-):** 10
- **Charge:** 0

**Magnesium**
- **Proton (+):** 12
- **Neutron (0):** 12
- **Electron (-):** 12
- **Charge:** 0

**Magnesium Ion**
- **Proton (+):** 12
- **Neutron (0):** 12
- **Electron (-):** 10
- **Charge:** -2

**Will Mg Gain or Lose e⁻?**

**Summary:** Mg in its ion form gains 2 electrons to achieve a noble gas configuration, becoming Mg²⁺.
General Rule For Ions: Atoms lose or gain electrons until they have the same number of electrons as the nearest noble gas atom.

Cation: A positive charged ion. Formed by elements on the left side of the periodic table.

Example: \( \text{Mg} \rightarrow \text{Mg}^{2+} + 2e^- \)

Note: Elements in groups 1 and 2 have charges usually equal to the group number.

Note: d-block and heavier elements have multiple possible charges.

Anion: A negatively charged ion. Formed by elements on the right side of the periodic table.

Example: \( \text{F} + e^- \rightarrow \text{F}^- \)

Note: Anions form a charge equal to their distance from the nearest noble gas.
Note: Zn, Ag, and Cd only for 1 ion, therefore, roman numerals are not used with compounds contain these ions. You must memorize the charge on these 3 transition metal ions.
Student Question

Which of the following statements is true?

a) Alkali metals tend to lose one electron to form 1\(^-\) ions.
b) Group 16 atoms tend to lose two electrons to form 2\(^+\) ions.
c) Halogens atoms tend to gain one electron to form 1\(^+\) ions.
d) More than one statement is true.
e) None of the previous statements are true.
**Student Question**

The correct number of protons, neutrons, and electrons (in that order) in the $^{53}{\text{V}}^{4+}$ ion is:

a) 23, 53, 19  
b) 23, 30, 23  
c) 23, 30, 19  
d) 19, 34, 19  
e) None of the above
Pure Substances

- **Pure Substance**: A substance composed of only one type of element or compound.

- **Element**: A substance that cannot be broken down into simpler substances.
  - **Atomic Element**: Elements that exist in nature with single atoms as their base unit.
  - **Molecular Elements**: Elements that exist as diatomic molecules in nature.

Examples: $H_2$, $N_2$, $O_2$, $F_2$, $Cl_2$, $Br_2$, and $I_2$
Pure Substances

- **We Know:** hydrogen(g) + chlorine(g) → HCl
- What would happen if H and Cl were monatomic?

1 L H
(x atoms)

Valve Closed

1 L Cl
(x atoms)

What would happen if H₂ and Cl₂ were diatomic?

1 L H₂
(x moles)

Valve Closed

1 L Cl₂
(x moles)
Pure Substances

- **Compound**: A substance composed of two or more elements in fixed, definite proportions.
  - **Molecular Compound**: electrically neutral molecules that are held together by sharing electrons. (usually contain only nonmetals)
  - **Ionic Compound**: Ions that are held together by electrostatic attraction which results in an overall electrically neutral compound. (usually contains metal and nonmetal)

<table>
<thead>
<tr>
<th>Molecular or Ionic</th>
<th>CS₂</th>
<th>CuO</th>
<th>KI</th>
</tr>
</thead>
</table>

**Pure Substances**

- **Chemical Formulas**: A collection of chemical symbols and subscripts that show the composition of a substance.

  - **Molecular Formula**: A combination of chemical symbols and subscripts showing the actual numbers of atoms of each element present in a molecule.
    
    **Note**: Used for molecular compounds only.

  - **Empirical Formula** *(formula unit)*: A chemical formula that shows the relative numbers of atoms of each element in a compound by using the simplest whole-number subscripts.
    
    **Note**: Used mainly for ionic compounds but can be used for molecular compounds.
Pure Substances

a) 
\[
\begin{array}{c}
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\end{array}
\]

Molecular Formula: \( \text{H}_2\text{C}_2\text{H}_5 \)
Empirical Formula: \( \text{CH}_3 \)

b) 
\[
\begin{array}{c}
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\end{array}
\]

Molecular Formula: \( \text{H}_2\text{O}_2\text{H} \)
Empirical Formula: \( \text{H}_2\text{O} \)

c) 
\[
\begin{array}{c}
\text{Cl}^- \\
\text{Na}^+ \\
\text{Na}^+ \\
\text{Cl}^- \\
\text{Cl}^- \\
\text{Na}^+ \\
\end{array}
\]

Empirical Formula: \( \text{NaCl} \)


d) 
\[
\begin{array}{c}
\text{OH}^- \\
\text{Ca}^{2+} \\
\text{OH}^- \\
\end{array}
\]

Empirical Formula: \( \text{Ca(OH)}_2 \)
Pure Substances

- **Molecular Formulas: Molecular Compounds**
  - **Subscripts** show number of atoms (or group of atoms) in molecule
  - **Parenthesize** show groups of atoms bonded together.
  - **Number before formula** shows number of molecules.

- **Empirical Formulas: Ionic Compounds**
  - **Subscripts** show number of atoms (or polyatomic ions) in crystal
  - **Parenthesize** show groups of polyatomic ions.
  - **Number before formula** shows larger crystal.
  - **Dot** shows loosely bound species

Example: \( \text{CuSO}_4 \cdot 5\text{H}_2\text{O} \)
General rule for writing chemical formulas

- Write the most metallic element first.

**Note:** Element nearest to bottom left hand corner of the periodic table or closest to Fr.

**Note:** This rule does not hold true for organic chemistry (compounds containing carbon).
**Pure Substances**

Ionic Compound = Cation + Anion = Neutral Species

<table>
<thead>
<tr>
<th>Atoms</th>
<th>Cation</th>
<th>Anion</th>
<th>Formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>Potassium</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Iodine</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Aluminum</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Chlorine</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Oxygen</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Magnesium</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Oxygen</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Boron</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
Ionic Cations

- **Rule 1**: Monatomic elements that only form 1 cation have the same name as the element. (Type I)
  
  *Example: Na\(^+\) sodium ion*

- **Rule 2**: For monatomic elements that form more than one cation, the oxidation number (the charge on the cation) is written in Roman numerals in a parentheses following the name. (Type II)
  
  *Note: This rule is necessary for most transition metals (exceptions: Ag\(^+\), Zn\(^2+\), Cd\(^{2+}\)).*

  *Examples: Fe\(^{2+}\) iron(II) ion*
Naming Compounds

Ionic Anions

- **Monatomic**: Monatomic anions are named by adding the suffix –ide to the “stem” of the name.

  *Example*: $\text{Cl}^-$ chloride ion

Ionic Compounds

- Ionic compounds are named by combining the cation name + the anion name.

  *Note*: The word ion does not appear in the compound name.

  *Examples*: $\text{K}^+$ potassium ion
  $\text{Cl}^-$ chloride ion
  $\text{KCl}$ potassium chloride
Naming Compounds

Chapter 2: Atoms, Molecules, and Ions
Naming Compounds

- NH$_4^+$ ammonium
- NO$_2^-$ nitrite
- NO$_3^-$ nitrate
- SO$_3^{2-}$ sulfite
- SO$_4^{2-}$ sulfate
- HSO$_4^-$ hydrogen sulfate
- OH$^-$ hydroxide
- CN$^-$ cyanide
- PO$_4^{3-}$ phosphate
- HPO$_4^{2-}$ hydrogen phosphate
- H$_2$PO$_4^-$ dihydrogen phosphate
- CO$_3^{2-}$ carbonate
- HCO$_3^-$ hydrogen carbonate
- ClO$^-$ hypochlorite
- ClO$_2^-$ chlorite
- ClO$_3^-$ chlorate
- ClO$_4^-$ perchlorate
- C$_2$H$_3$O$_2^-$ acetate
- MnO$_4^-$ permanganate
- Cr$_2$O$_7^{2-}$ dichromate
- CrO$_4^{2-}$ chromate
- O$_2^{2-}$ peroxide

If a polyatomic ion is present, substitute the name of the polyatomic ion for either the cation or the anion.
Polyatomic Anions Containing Oxygen (Oxoanions)

**Rule 1:** If only one oxoanion of an element exists, its name is formed by adding the suffix –ate to the stem of the name of the element.

*Example:* $\text{CO}_3^{2-}$ carbonate ion

**Rule 2:** For elements that can form two types of oxoanions, the ion with the larger number of oxygen atoms is given the suffix –ate and that with the smaller number of oxygen atoms is given the suffix –ite.

*Example:* $\text{NO}_2^-$ nitrite ion
$\text{NO}_3^-$ nitrate ion
Rule 3: For elements that can form more than two kinds of oxoanions, the oxoanion with the smallest number of oxygen atoms is formed by adding the prefix hypo- to the -ite form of the name.

Example: ClO\textsuperscript{-} hypochlorite ion

The oxoanion with the largest number of oxygen atoms is formed by adding the prefix per- to the -ate form of the name.

Example: ClO\textsubscript{4}\textsuperscript{-} perchlorate ion
Inorganic Molecular Compounds

**Rule 1:** Put a Greek prefix in front of the name to indicate the number of each type of atoms present.

*Note:* Mono is never used for 1st atom.

**Example:** NO  
Nitrogen monoxide

<table>
<thead>
<tr>
<th>Prefix</th>
<th>Number</th>
</tr>
</thead>
<tbody>
<tr>
<td>mono-</td>
<td>1</td>
</tr>
<tr>
<td>di-</td>
<td>2</td>
</tr>
<tr>
<td>tri-</td>
<td>3</td>
</tr>
<tr>
<td>tetra-</td>
<td>4</td>
</tr>
<tr>
<td>penta-</td>
<td>5</td>
</tr>
<tr>
<td>hexa-</td>
<td>6</td>
</tr>
<tr>
<td>hepta-</td>
<td>7</td>
</tr>
<tr>
<td>octa-</td>
<td>8</td>
</tr>
<tr>
<td>nona-</td>
<td>9</td>
</tr>
<tr>
<td>deca-</td>
<td>10</td>
</tr>
<tr>
<td>undeca-</td>
<td>11</td>
</tr>
<tr>
<td>dodeca-</td>
<td>12</td>
</tr>
</tbody>
</table>

**Rule 2:** For binary molecular compounds, name the element that occurs further to the top right of the periodic table second with its ending changed to –ide.

*Note:* The final a or o of the prefix is usually dropped if the element begins with a vowel.
Naming Compounds

Example: \( \text{P}_2\text{O}_5 \)
diphosphorus pentoxide
Naming Compounds

- **Acid (Arrhenius)**: A compound that contains hydrogen and reacts with water to form hydrogen ions.
  
  *Example*: HCl(aq)

- **Aqueous Solution**: A solution in which the solvent is water.

**Acids**

- **Rule 1**: Binary acids are named by adding the prefix hydro- and changing the ending of the name of the second element to -ic acid.

  *Note*: If you have HCl(g) (not dissolved in water) it is called hydrogen chloride but once it is dissolved in water, HCl(aq), it is called hydrochloric acid. Notice that the compound is not hydrogen monochloride; this is because hydrogen can only form one bond, therefore, the chemical formula can be determine without Greek prefixes.
Naming Compounds

Oxoacids (acids containing oxygen)

- Oxoacids are the parents of oxoanion

Examples: \( \text{H}_2\text{SO}_4 \rightarrow 2\text{H}^+ + \text{SO}_4^{2-} \)
\( \text{H}_2\text{SO}_3 \rightarrow 2\text{H}^+ + \text{SO}_3^{2-} \)

- **Rule 1:** If the oxoanion ends in –ate change –ate to –ic acid.

- **Rule 2:** If the oxoanion ends in –ite change –ite to -ous acid

Examples: \( \text{H}_2\text{SO}_4(aq) \) sulfuric acid \( \text{SO}_4^{2-} \) sulfate ion
\( \text{H}_2\text{SO}_3(aq) \) sulfurous acid \( \text{SO}_3^{2-} \) sulfite ion

**Note:** If the oxoanion starts with per or hypo leave the prefix in the acid name.
Take Away From Chapter 2

**Big Idea:** All matter is made of atoms. Atoms contain three fundamental particles. The number of protons in an atom defines the type of atoms. Atoms join together to form larger compounds. These compounds belong to one of two classes: molecular or ionic compounds.

- **Atoms**
  - Know the three law that lead to the understanding that all matter is made from atoms (90,91,94)
    - Law of conservation of mass
    - Law of definite proportions
    - Law of multiple proportions
  - Understand the historical experiments that identified the fundamental participles in an atom
    - J.J Thomas/cathode ray tube (existence of e⁻)
    - Millikan/oil drop (charge of e⁻)
    - Rutherford/ Pt foil experiment (size and location of nucleus)

Numbers correspond to end of chapter questions.
Take Away From Chapter 2

- **Atoms** (continued)
  - Know what a atom looks like and the name and symbol of elements 1-54. (28)
  - Know that the atomic number \( Z \) is the number of protons and that the mass number \( A \) is the number of protons and neutrons and how to indicate the isotope using short hand notation. (42,43,44,45)

- **Periodic Table**
  - Know the names and locations of major groups (ex: halogens group 17). (37)
  - Be able to determine if an element is a metal, nonmetal, or metalloid. (34,36)

- **Ions**
  - Be able to determine what ion elements form (47,48)

Numbers correspond to end of chapter questions.
Take Away From Chapter 2

- **Pure Substances**
  - Know which elements are diatomics (19,21)
    - H₂, N₂, O₂, F₂, Cl₂, Br₂, and I₂
  - Be able to distinguish between molecular (nonmetals) and ionic (metal + nonmetal) compounds and know the type of bonds that hold molecular (covalent) and ionic (ionic) compounds together (32,65)
  - Be able to determine the number and type of atoms in a given chemical formula
  - Be able to determine ionic formula units (71,78)

- **Naming Compounds (54,55,56,57,58,59,60)**
  - Be able to name ionic compounds
    - Know the name of polyatomic ions (Table 2.5)
  - Be able to name molecular compounds (53)
    - Be able to name acids

Numbers correspond to end of chapter questions.