Chapter 2: Atoms, Molecules, and Ions

1. Atomic Structure
2. Atomic Number & Mass
3. Isotopes
4. Atomic Weight
5. Periodic Table
6. Molecular Formulas & Nomenclature
7. Molecular Compounds
8. Atoms, Molecules & Mole
9. Compound Formulas
10. Hydrated Compounds

Structure of The Atom
Modern View

Three Fundamental Particles

1. **Proton** (Positive Charged Particle)
2. **Neutron** (Neutral Particles)
3. **Electron** - Negative Charged Essentially Mass Less Particle, Occupies Majority of Atomic Space

**Nucleus** - Region of Atomic Mass Density, Contains Neutrons and Protons

Protons, Neutrons & Electrons: Historical Perspective

**Electricity**

Labeled 2 types of charge as Positive and Negative

Like Charges Repel
Opposite Charges Attract

Noted that charge is balanced, so for every [+] charge there exists a [-] charge

Types of Radioactive Processes

1. **Alpha Decay**
   - Heavy [+] Particle
2. **Beta Decay**
   - Light [-] Particle
3. **Gamma Decay**
   - Charge less
   - Release of Energy

Some History on The Discovery of the Atom

Discovery of The Electron:
J.J. Thomson and the Cathode Ray Tube

Calculated $e^-/m$

Some History on The Discovery of the Atom

Discovery of The Electron:
Robert Millikan (1911) Measured the Charge of the Electron

Charge of electron $= 1.6 \times 10^{-19}$C.
Some History on The Discovery of the Atom
Rutherford’s Gold Foil Experiment and the Discovery of the Nucleus
Believed in Plum Pudding Model
Wanted to see How Large Atoms are

![Image of alpha particles](Image)

2.1 Structure of The Atom
Modern View

<table>
<thead>
<tr>
<th>Particle</th>
<th>Relative Mass (g)</th>
<th>Relative Charge (amu)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Proton</td>
<td>$1.6726 \times 10^{-24}$</td>
<td>1.0073 +1</td>
</tr>
<tr>
<td>Neutron</td>
<td>$1.6749 \times 10^{-24}$</td>
<td>1.0087 0</td>
</tr>
<tr>
<td>Electron</td>
<td>$9.1094 \times 10^{-28}$</td>
<td>0.0005486 -1</td>
</tr>
</tbody>
</table>

Atomic Number & Mass Number

The Mass of an Atom Can Be Determined From the Masses of the Neutrons and Protons in It’s Nucleus

Atomic Mass Unit (AMU)

$1\text{amu} = 1.66 \times 10^{-24}\text{g}$

This Is A Very Small Number

1 Amu = Mass of 1 Carbon Atom With 6 Neutrons & 6 Protons

2.3 Isotopes

The Identity of an Element Is Determined by the Number of Protons It Possesses

Isotopes - Elements With the Same Number of Protons but Different Numbers of Neutrons. Different Atoms of Same Element Can Have Different Masses.

**Isotopic Notation:**

- $A = \text{Mass Number (# of P and N)}$
- $Z = \text{Atomic Number (# of P)}$
- $X = \text{Elemental Symbol}$

Isotopes

Determine the Number of Protons, Neutrons and Electrons for the 3 Isotopes of Oxygen

16$^8$O:

17$^8$O:

18$^8$O:

Note: For Neutral Atoms, the # of Protons Equals the # of Electrons
Isotopes of Monatomic Ions
Note: For ions there is one less electron than proton for each [+] charge and one more electron than protons for each [−] charge.

Determine the Number of Protons, Neutrons and Electrons for the 2 Ions:

\[ ^{16}_{8}O^{-2} : \]
\[ ^{55}_{25}Mn^{+7} \]

Radioactive Isotopes
Many Isotopes Are Unstable and Undergo Radioactive Decay.

If the Number of Protons Changes During a Radioactive Process, the Identity of the Element Changes.

We Will Cover These in Chapter 23 of the Text

2.4 Atomic Weight

Given in the Periodic Table.

Why Is the Atomic Weight of a Carbon = 12.01 Amu?

Isotope Abundance

Natural Samples of Pure Elements Contain a Mixture of that Elements Isotopes

\[ \text{Isotopic Abundance} = \frac{\# \text{ of Atoms of Given Isotope}(100)}{\text{Total Atoms of All Isotopes}} \]

Class Problem

Magnesium has 3 Isotopes with the following masses and % Abundances,
23.985 amu (78.99%),
24.986 amu (10.00%)
25.983 amu (11.01%)

Determine the average atomic weight of Magnesium

Class Problem

Lithium has has only 2 isotopes, \(^{6}\text{Li}\) (with mass of 6.0151) and \(^{7}\text{Li}\) (with mass of 7.0160)

\[ \text{What is the \% Abundance for Each Isotope?} \]
Let
\[ X = \text{fraction of } ^{6}\text{Li} \]
\[ Y = \text{fraction of } ^{7}\text{Li} \]

\[ X(6.0151) + Y(7.0160) = 6.941 \]

2 Unknowns Requires 2 Equations
\[ X + Y = 1 \quad \Rightarrow \quad Y = 1 - X \]

---

### Periodic Table

Be Able to Identify the Following Groups or Families From the Periodic Table

1. Alkali Metals (Group 1)
2. Alkaline Earth Metals (Group 2)
3. Halogens (Group 7)
4. Noble Gases (Group 8)
5. Transition Metals
6. Lanthanides
7. Actinides

---

### 2.5 Periodic Table

**Physical Properties of Metals**

1. **Conduction of Heat and Electricity**
2. **Malleability** (can be hammered into thin sheets)
3. **Ductility** (can be pulled into wires)
4. **Lustrous** (shiny) appearance

---

### Metals, Nonmetals & Metalloids

**Nonmetals**

- Poor Conductors
- Nonlusterous
- Nonmalleable
- Solids are brittle
- Some are brightly colored
Natural States of The Elements

**Liquids** - Hg, Br₂

**Gasses** - Noble Gasses and lighter diatomics, H₂, N₂, O₂, F₂, Cl₂

**Solids** - All other elements

Elements Which Often Appear in Pure Form

A) Noble Metals (Ag, Au, & Pt)
B) Noble Gases (He, Ne, Ar, Kr Xe & Rn)

“Noble” Implies “Unreactive”

All elements with Atomic # > 83 are Radioactive

---

7 Diatomics

Know the Diatomic’s and Their Natural States

<table>
<thead>
<tr>
<th>Element</th>
<th>Formula</th>
<th>State</th>
<th>Color</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydrogen</td>
<td>H₂</td>
<td>gas</td>
<td>colorless</td>
</tr>
<tr>
<td>Nitrogen</td>
<td>N₂</td>
<td>gas</td>
<td>colorless</td>
</tr>
<tr>
<td>Oxygen</td>
<td>O₂</td>
<td>gas</td>
<td>pale blue</td>
</tr>
<tr>
<td>Fluorine</td>
<td>F₂</td>
<td>gas</td>
<td>pale yellow</td>
</tr>
<tr>
<td>Chlorine</td>
<td>Cl₂</td>
<td>gas</td>
<td>pale green</td>
</tr>
<tr>
<td>Bromine</td>
<td>Br₂</td>
<td>Liquid</td>
<td>red/brown</td>
</tr>
<tr>
<td>Iodine</td>
<td>I₂</td>
<td>Solid</td>
<td>purple</td>
</tr>
</tbody>
</table>

---

Group 1A: Alkali Metals

H, Li, Na, K, Rb, Cs & Fr

Very Reactive – Do Not Exist as Pure Elements in Nature

Form Oxides of Formula A₂O

---

Group 2A: Alkali Earth Metals

Be, Mg, Ca, Sr, Ba & Ra

Reactive – Form Alkaline Solutions (Basic)

Form Oxides of Formula AO

---

Group 3A

B, Al, Ga, In, Tl

Boron comes from Borax

Form Oxides of Formula A₂O₃
### Group 4A

**C, Si, Ge, Sn, Pb**

- **C** - non metal
- **Si, Ge** - metalloids
- **Sn, Pb** - metals

**Carbon forms Allotropes**

(Different Forms of the Same Element)

- Diamond
- Graphite
- Buckminsterfullerene

### Group 5A

**N, P As, Sb, Bi**

- **N, P** - Non metals (essential to life, N is most abundant element in atmosphere)
- **As, Sb** - metalloids
- **Bi** - metal, heaviest non radioactive element

### Group 6A

**O, S, Se, Te, Po**

- **O, S, Se** - nonmetals
- **Te** - metalloid
- **Po** - Radioactive metal

**Chalcogens**

**Chalk Formers**

### Group 7A

**F, Cl, Br, I, At**

- **Halogen**
- **Salt Formers**

**F, Cl, Br, I** - nonmetal diatomics

**At** - often classified as radioactive metalloid

### Group 8A

**He, Ne, Ar, Kr, Xe**

- **Noble Gases**

“**Inert Gases**” – do not tend to form compounds
4 Types of Chemical Bonds

1. **Ionic Bonds** - Ionic compounds between metals and nonmetals consisting of [+] cations and [-] anions.

2. **Covalent Bonds** - Classical “molecules” where valence electrons are shared between two nonmetals.

3. **Polar Covalent Bonds** - Covalent bonds with ionic character in that the electrons are not equally shared.

4. **Metallic Bonds** - Pure metals and alloys where delocalized free electrons hold together the positive nuclei.

2.6 Molecules & Compounds

- Pure Substance Composed of More Than 1 Atom

**Formulas** Describe the Elemental Composition

- Common Names Like Baking Soda or Sugar Tell Us Nothing About Their Composition

- Their Formulas Provide Information on Their Chemical Composition

Chemical Formulas

Sugar = C_{12}H_{22}O_{11}
Baking Soda = NaHCO_{3}

This tells us that:
- one molecule of sugar has 12 carbon atoms, 22 hydrogen atoms and 11 oxygen atoms
- Baking Soda has 3 atoms of Oxygen and 1 atom each of Sodium, Hydrogen and Carbon

Molecular Formulas

Molecular Formulas - only tell of the elemental composition

- both ethanol and diethyl ether have the same molecular formula, C_{2}H_{6}O

- Yet they are different

Structural Formula

Ethanol = CH_{3}CH_{2}OH
Dimethyl Ether = CH_{3}OCH_{3}

H H
H-C-C-O-H H-C-O-C-H
H H bonds H H

2.7 Ions & Nomenclature

Charged Atoms Occur When an Atom Losses or Gains **Valence** Electron(s)

Two Types:

- **Anions** - Negative Ions (Gain Electrons)
- **Cations** - Positive Ions (Lose Electrons)

Note: Charge Is Indicated by a +/- Post Superscript
Ions

Ionic Charge and the Periodic Table:

<table>
<thead>
<tr>
<th>Cations</th>
<th>Anions</th>
</tr>
</thead>
<tbody>
<tr>
<td>+1 Alkali Metals, H, Ag</td>
<td>-1 Halogens, Hydrogen</td>
</tr>
<tr>
<td>+2 Alkaline Earths, Zn, Cd</td>
<td>-2 Group VI Nonmetals</td>
</tr>
<tr>
<td>+3 Aluminum</td>
<td>-3 Group V Nonmetals</td>
</tr>
</tbody>
</table>

Noble Gasses Do Not Tend to Form Ions

Note - Monatomic Ions Tend to Have Charge Required to Become Isoelectronic With Nearest Nobel Gas (Have Same # of Electrons As the Noble Gas)

Polyatomic Ions

-Charged Covalent Compounds

\[
\begin{align*}
\text{H} & \quad \text{O} \\
\text{H-N-H} & \quad \text{O-S-O} \\
\text{H} & \quad \text{O}
\end{align*}
\]

\[\text{NH}_4^+ \quad \text{SO}_4^{2-}\]

Formula Unit

-Ionic Compounds Are Identified by Their Formula Unit
-Formula Unit Represents the Lowest Whole # Ratio of Ions in the Crystal Lattice

Nomenclature

1. Ionic Compounds (between metals and nonmetals)
2. Covalent Compounds (between two nonmetals)
3. Compound with Polyatomic Ions
4. Acids
5. Compound Formulas from Names

Binary Compounds

Binary Compounds - Contain two types of atoms, can be ionic or covalent

Examples:

\[\text{H}_2\text{O, CO, CO}_2, \text{ CoCl}_2, \text{ FeO, Fe}_2\text{O}_3, \text{ NaCl}\]

Binary Covalent  Binary Ionic
**Binary Ionic Compounds**

Form Between a Metal and a Nonmetal,
- Have Monatomic Ions
- Form Crystal Lattice Type Solid Structures
- Oxidation # of a Monatomic Ion Is It's Charge

**Cation** - [+] ion, (*Metals Form Cations*)
Type 1 Metal (forms only 1 cation) Na⁺,
Type 2 Metal (forms 2 or more cations) Fe²⁺, Fe³⁺,...

**Anion** - [-] ion, (*Nonmetals Form Anions*)
ex: Cl⁻, O²⁻, N³⁻,...

**Binary Ionic Compounds**

**Type I Metals** - have single charge in all ionic compounds (Invariant Oxidation #)

Know Following Invariant Oxidation #’s:
+1 Alkali Metals, H, Ag
-1 Halogens, Hydrogen
+2 Alkaline Earths, Zn, Cd
-2 Group VI Nonmetals
+3 Aluminum
-3 Group V Nonmetals

Note - “A” Group Monatomic Ions Tend to Have Charge Required to Become Isoelectronic With Nearest Noble Gas (Have Same # of Electrons As the Noble Gas)

**Binary Ionic Compounds**

Compounds with Type I Metals

**Rules:**
1. Name Metal First
2. Name Nonmetal 2nd With -ide Suffix

Ex: Table Salt, NaCl = Sodium Chloride
BaCl₂ = Barium Chloride

**Binary Ionic Compounds**

**Ionic Formulas from Names**

**Rules:**
1. Ionic Formula Must Be Neutral (Principal of Charge Balance)
2. Subscripts After Elemental Symbol Indicate # of That Element in Formula
3. Correct Formula Indicates Lowest Whole # Ratio of Anions to Cations

Magnesium Chloride = MgCl₂ not Mg₂Cl₄

**Binary Ionic Compounds**

**Ionic Formulas from Names**

Trick: Set # of Anions = Charge of Cation
and # of Cations = |Charge of Anion|
Beware That This Is the Lowest Whole # Ratio

Consider Aluminum Oxide,
Charge of Aluminum = +3 and Oxygen = -2.

\[
\text{Al}_2\text{O}_3
\]

\[
(2\times+3)+(2\times-3)=0
\]

<table>
<thead>
<tr>
<th></th>
<th>-1</th>
<th>-2</th>
<th>-3</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>Cl</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Na</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>2</td>
<td>Cl</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Na</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>3</td>
<td>Cl</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
Binary Ionic Compounds

**Type II Metals** - have multiple charges in ionic compounds (Variable Oxidation #'s)

<table>
<thead>
<tr>
<th>Know Following Variable Oxidation State Ions</th>
</tr>
</thead>
<tbody>
<tr>
<td>Fe&lt;sup&gt;3+&lt;/sup&gt; Iron(III) - ferric</td>
</tr>
<tr>
<td>Fe&lt;sup&gt;2+&lt;/sup&gt; Iron(II) - ferrous</td>
</tr>
<tr>
<td>Cu&lt;sup&gt;2+&lt;/sup&gt; Copper(II) - cupric</td>
</tr>
<tr>
<td>Cu&lt;sup&gt;+&lt;/sup&gt; Copper(I) - cuprous</td>
</tr>
<tr>
<td>Co&lt;sup&gt;3+&lt;/sup&gt; Cobalt(III) - cobaltic</td>
</tr>
<tr>
<td>Co&lt;sup&gt;2+&lt;/sup&gt; Cobalt(II) - cobaltous</td>
</tr>
</tbody>
</table>

**Binary Ionic Compounds**

**Type II Metals**

**Rules:**

1. Name **Metal First**
2. Indicate **Metal’s Charge** (Oxidation State) in **Roman Numerals**
   
   FeCl<sub>2</sub> = Iron(II)Chloride
   FeCl<sub>3</sub> = Iron(III)Chloride

**Naming Covalent Compounds**

**Compounds Between Nonmetals** Do Not Form Ions but Share Electrons (Forming Covalent Bonds)

- These Are the Classical “**Molecules**” (in Contrast to Ionic Crystal Lattices)
- Use **Greek Prefixes** to Indicate Number of Each Type of Atom in Molecules
- Write “**More Metallic**” Nonmetal First
- If More Than 2 Atoms, Write in Order of Connectivity (SCN vs. SNC)

**Greek Prefixes**

| 1 | mono- |
| 2 | di-   |
| 3 | tri-  |
| 4 | tetra- |
| 5 | penta- |
| 6 | hexa- |
| 7 | hepta-|
| 8 | octa- |
| 9 | nona- |
| 10| deca- |

Note: often do not use prefix mono-

**Naming Covalent Compounds**

**Group Problems: Name or write the following formula’s**

CCl<sub>4</sub>
SO<sub>3</sub>
Dinitrogen trioxide
Diphosphorous pentoxide

**Polyatomic Ions**

Many Ionic Compounds Have Covalently Bonded Polyatomic Ions

Rules Are Essentially Same, but Use the Name of the Polyatomic Ion, That Is;

**Cation First** (Charge If Variable) **Anion Last**
### Polyatomic Ions

<table>
<thead>
<tr>
<th>Ion</th>
<th>Name</th>
<th>Formula</th>
<th>Polyatomic Ion</th>
</tr>
</thead>
<tbody>
<tr>
<td>NH₄⁺</td>
<td>Ammonium</td>
<td>CH₄</td>
<td></td>
</tr>
<tr>
<td>CH₃COO⁻</td>
<td>Acetate</td>
<td>CH₃COO⁻</td>
<td></td>
</tr>
<tr>
<td>NO₃⁻</td>
<td>Nitrite</td>
<td>NO₃⁻</td>
<td></td>
</tr>
<tr>
<td>NO₂⁻</td>
<td>Nitrate</td>
<td>NO₂⁻</td>
<td></td>
</tr>
<tr>
<td>OH⁻</td>
<td>Hydroxide</td>
<td>OH⁻</td>
<td></td>
</tr>
<tr>
<td>ClO₄⁻</td>
<td>Perchlorate</td>
<td>ClO₄⁻</td>
<td></td>
</tr>
<tr>
<td>ClO₃⁻</td>
<td>Chlorate</td>
<td>ClO₃⁻</td>
<td></td>
</tr>
<tr>
<td>ClO₂⁻</td>
<td>Chlorite</td>
<td>ClO₂⁻</td>
<td></td>
</tr>
<tr>
<td>OH⁻</td>
<td>Hydronate</td>
<td>OH⁻</td>
<td></td>
</tr>
<tr>
<td>HSO₄⁻</td>
<td>Perbeiginate</td>
<td>HSO₄⁻</td>
<td></td>
</tr>
<tr>
<td>HSO₃⁻</td>
<td>Perbrominate</td>
<td>HSO₃⁻</td>
<td></td>
</tr>
<tr>
<td>HSO₂⁻</td>
<td>Perbromit</td>
<td>HSO₂⁻</td>
<td></td>
</tr>
<tr>
<td>HO⁻</td>
<td>Hypobromite</td>
<td>HO⁻</td>
<td></td>
</tr>
<tr>
<td>H₃PO₄⁻</td>
<td>Perhydrol</td>
<td>H₃PO₄⁻</td>
<td></td>
</tr>
<tr>
<td>H₃PO₃⁻</td>
<td>Perhydrolate</td>
<td>H₃PO₃⁻</td>
<td></td>
</tr>
<tr>
<td>H₃PO₂⁻</td>
<td>Perhydrite</td>
<td>H₃PO₂⁻</td>
<td></td>
</tr>
<tr>
<td>HO⁻</td>
<td>Hypochlorite</td>
<td>HO⁻</td>
<td></td>
</tr>
<tr>
<td>H₃AsO₄⁻</td>
<td>Perarsenate</td>
<td>H₃AsO₄⁻</td>
<td></td>
</tr>
<tr>
<td>H₃AsO₃⁻</td>
<td>Perarsenite</td>
<td>H₃AsO₃⁻</td>
<td></td>
</tr>
<tr>
<td>H₃AsO₂⁻</td>
<td>Perarsit</td>
<td>H₃AsO₂⁻</td>
<td></td>
</tr>
<tr>
<td>HO⁻</td>
<td>Hypochlorite</td>
<td>HO⁻</td>
<td></td>
</tr>
</tbody>
</table>

### Polyatomic Ions

Tip for Remembering Charges - Many **Oxyanions** (Polyatomic Ions With Multiple Oxygens) Have the Same Charge As the Monatomic Nonoxygen Ion

- Sulfide = -2 As Does Sulfate & Sulfite
- Chloride = -1 As Does Perchlorate, Chlorate, Chlorite & Hypochlorite
- Phosphate = -3 Does Phosphate & Phosphate

**Note, Nitrogen Is an Exception:**

- Nitride = -3, Nitrate & Nitrite = -1

### Acid Nomenclature

Compound which can release H⁺ in water are called Acids. (aqueous, aq., means dissolved in water)

Nomenclature is related to that of Ionic Compounds

- If salt is binary, it forms binary acids:
  - **ide** => **hydro(anions elemental name)**-ic acid

If salt has oxyanions it forms oxyacids with endings:

- **ate** => **ic acid**
- **ite** => **ous acid**

### Sulfur’s Sodium Salts and Acids

- **Na₂S** - Sodium Sulfide
  - H₂S(aq) - **Hydrosulfuric Acid**
    - Note: H₂S(g) - Hydrogen Sulfide (gas)

- **Na₂SO₄** - Sodium Sulfate
  - H₂SO₄(aq) - Sulfuric Acid

- **Na₂SO₃** - Sodium Sulfite
  - H₂SO₃(aq) - Sulfurous Acid

### Acid Nomenclature

Name or give Formulas for Following Acids:

- Acetic Acid (vinegar) -
- Chloric Acid -
- Hydrochloric Acid -
  - H₃PO₄ -
  - HCN -
Acid Salts

-Ions with more than one negative site can combine with more than one cation.

- Acid salts result when one cation is a proton

Let's look at all of the compounds that can form from phosphate

**Phosphate** = \( \text{PO}_4^{3-} \)

- \( \text{Na}_3\text{PO}_4 \) = Sodium Phosphate
- \( \text{Na}_2\text{HPO}_4 \) = Sodium Hydrogen Phosphate
- \( \text{NaH}_2\text{PO}_4 \) = Sodium Dihydrogen Phosphate
- \( \text{H}_3\text{PO}_4 \) = Phosphoric Acid

Name the compounds that can form from hydrogen, potassium and phthlate; and identify them as salts, acids or acid salts

Hydrated Salts

-Salts which incorporate water into their crystal structure are hydrated salts

-Include the water in the compound formula

-Greek prefixes indicate the # of waters of hydration

\( \text{CaCl}_2\,6\text{H}_2\text{O} \) = Calcium Chloride hexahydrate

What is the name of \( \text{CuSO}_4\cdot5\,\text{H}_2\text{O} \)?

**3.3 Molar Mass**

**Molar Mass** - Is the Mass (in Grams) of One Mole of a Substance (often called molecular weight or formula weight)

**What is the Molar Mass of Sulfate?**

1 mole \( \text{SO}_4^{2-} \): has 1 mole S and 4 moles O

Molar Mass

How many moles are in 5.86g of AgCl?

1 mol AgCl = 143.3g, i.e.,

\[ MM(\text{AgCl})=143.3\text{g/mol} \]

Molar Mass

Calculate the Molar Mass of:

1. Benzene (\( \text{C}_6\text{H}_6 \))
2. Acetylene (\( \text{C}_2\text{H}_2 \))
3. Water
4. Hydrogen Peroxide
Molar Mass

What is the Molar Mass for Aluminum Sulfate?

\[ \text{Al}_2(\text{SO}_4)_3 \] has 2 Al Cations and 3 Sulfate Anions

Note: We used the Molar Mass of Sulfate which we determined in the Previous Problem

Molar Mass

Calculate Molar Mass for:
1. Sodium Bicarbonate
2. Ammonium Nitrate
3. Potassium Perchlorate
4. Silver Chloride

Percent Composition

What is the Mass Fraction of an Element in a Compound?

\[
\text{Mass Fraction of a given element} = \frac{\text{Mass of element in compound}}{\text{Mass of compound}}
\]

What is the Mass Percent of an Element in a Compound?

\[
\text{Mass Percent of a given element} = \text{Mass Fraction of a given element} \times 100
\]

Percent Composition

1. What Is the Sum of the Mass Percents of All the Elements in a Compound?
2. Does the Percent Composition Change as the Quantity of Substance Changes?

What is the mass percent of H, C & O in Acetic Acid (Vinegar)

\[ \text{CH}_3\text{COOH} \]

Percent Composition

Consider 1 mole \( \text{CH}_3\text{COOH} \)

\[
\begin{align*}
M_{\text{Carbon}} &= 2(12.01) = 24.02g \\
M_{\text{Oxygen}} &= 2(16.00) = 32.00g \\
M_{\text{Hydrogen}} &= 4(1.01) = 4.04g \\
M_{\text{Acetic Acid}} &= 60.06g
\end{align*}
\]

How Do We Get the Percent Composition From the Above Data?
### Percent Composition Problems

1. Determine the % Composition of the Sulfate Ion.
2. Determine the % Composition of Aluminum Sulfate.
3. Determine the Percent Composition of Benzene ($C_6H_6$).
4. Determine the Percent Composition of Acetylene ($C_2H_2$).

### Compound Formulas

2. **Molecular Formula** - the Actual Ratio of the Various Elements in a Compound or Molecule.

Note: Benzene ($C_6H_6$) & Acetylene ($C_2H_2$) Have Different Molecular Formulas, but the Same Empirical Formula (CH).

### Empirical Formulas

- Empirical data is based on observation and experiment, not theory.
- Opposite Calculation to Percent Composition.
  - For Empirical Formulas, You Know the Masses of Each Element From Experimental Data, So Determine Simplest Formula.
  - In % Composition, You Start Knowing the Formula, and Determine the Mass %.

### Empirical Formulas

Determine the Empirical Formula of Aspirin, Which Was Analyzed and Found to Have a Mass Percent Composition of 60.0%C, 4.48%H and 35.5%O.

### Empirical Formulas

1. Obtain Mass of Each Element (in Grams) If given % Composition, assume 100 g of substance.
2. Calculate # of Moles of Each Element Present From Masses and Atomic Weights.
3. Divide # of Moles by the Moles of Least Present Element (Forcing It to 1).
4. Multiple the Results of Step 3 by Smallest Integer Which Will Convert Them All to Whole Numbers.

### Empirical Formulas

1. Assume 100 grams of sample, giving
   \[ M_C = \quad M_H = \quad M_O = \]
2. Calculate Moles Present
   \[
   \text{moles } C = \\
   \text{moles } H = \\
   \text{moles } O = 
   \]
Empirical Formulas

3. Divide By
   Smallest Mole
   Fraction (2.22)

4. Multiply by smallest
   Integer Forcing all
   results from step 3 to
   be whole numbers

Molecular Formulas

In Order to Determine the Molecular
Formula, We Need to Know the
Molar Mass in Addition to the
Empirical Formula

The Molecular Formula will be an Integral
Multiple of the Empirical Formula

For an Ionic Compound, the Formula Weight
Corresponds to the Empirical Formula

---

Determine the molecular formula of a compound
containing H, O & C. A 50.00 g sample has 23.88 g
of both carbon & oxygen and the compound has a
molar mass of 603 g/mol.

1. Determine moles in sample
   note: mass hydrogen = total mass - mass oxygen - mass carbon
   = 50.00 g - 2(23.88 g) = 2.24 g

Step 2: Divide by Smallest number
   Trick: Convert to Fractions

Tips in converting decimals to fractions

\[ \frac{1}{X} \text{ decimal expression} \Rightarrow X = \frac{1}{\text{decimal expression}} \]

Know:

\[
\begin{align*}
0.50 &= 1/2 \\
0.33 &= 1/3 \\
0.25 &= 1/3 \\
0.20 &= 1/5 \\
\end{align*}
\]

Step 3: Find lowest whole number ratio by
   factoring out the denominators

Step 4: Calculate Molecular Formula
   from Empirical Formula Using Masses

\[
\begin{align*}
EF &= \text{Empirical Formula} \\
EM &= \text{Empirical Mass} \\
MF &= \text{Molecular Formula} \\
MM &= \text{Molecular Mass} \\
MF &= n(EF) & MM &= n(EM), n=1,2,3\ldots \\
EM(C_8H_9O_6) &= 201 g/mol \quad MM=603 g/mol \\
\frac{MM}{EM} &= n \\
MF &= n(EF)=
\end{align*}
\]
### Empirical Formula of Formic Acid from Combustion Data

<table>
<thead>
<tr>
<th>Sample (C₅H₆O₅)</th>
<th>furnace</th>
<th>H₂O absorber</th>
<th>CO₂ absorber</th>
<th>excess oxygen</th>
</tr>
</thead>
<tbody>
<tr>
<td>2.4541g formic acid</td>
<td></td>
<td>0.9606g</td>
<td>2.3482g</td>
<td></td>
</tr>
</tbody>
</table>

Step 1: Determine Mass of each element

\[
\begin{align*}
m_H &= 0.9606g(2gH/18gH_2O) = 0.1067g H \\
m_C &= 2.3482g(12gC/44gCO_2) = 0.6404g C \\
m_O &= 2.4541g - 0.1067g - 0.6404g = 1.7071g O
\end{align*}
\]

Step 2: Determine Mole of Each Element

Step 3: Divide by Smallest Number

### 2.11 Hydrated Salts

**Hydrated salts** absorb water into the crystal lattice—this is the “water of hydrations”

- Stoichiometric Proportions (water of hydration is of integer proportions)

**Anhydrous Salt**—“Without Water”, Form of Salt without water of hydration, forms a different type of crystal

### Hydrated Salts

Sodium Carbonate Forms a Hydrated Salt

\[Na_2SO_4\cdot xH_2O\]

What is the Formula if a 0.767g sample was dried to a constant weight of 0.284 g?

### Hydrated Salts

\[
M_{hydrated \ salt} = M_{anhydrous \ salt} + M_{water}
\]

\[
M_{water} = M_{hydrated \ salt} - M_{anhydrous \ salt}
\]

\[
M_{water} = 0.767g - 0.284g = 0.483g
\]

Dividing moles water and moles Sodium carbonate by species with least number of moles (0.00268 moles Na₂CO₃):

\[Na_2CO_3\cdot 10H_2O\]