Chapter I

Homogeneous - uniform throughout
Heterogeneous - not uniform throughout

Pure substance

Element

Compound

Physical change - changes physical appearance, state of matter, but not composition

Chemical change - chemically distinct substances example: lighting something on fire

Intensive property - do not depend on amount of sample density, color, temperature, boiling point

Extensive property - does depend on amount mass, volume

Density = \[
\frac{\text{mass}}{\text{volume}}
\]

Temperature

\[
\begin{align*}
K &= ^\circ C + 273.15 \\
^\circ C &= \frac{5}{9} (^\circ F - 32) \\
^\circ F &= \frac{9}{5} (^\circ C) + 32
\end{align*}
\]

Precision - how closely individual measurements agree with one another

Accuracy - how closely individual measurements agree with correct value

Significant figures - know the rules!

Subtraction/Addition - same # of decimal places as measurement with fewest decimal places

Multiplication/Division - same # of sig figures as measurement with the fewest sig figures
Chapter 2

Dalton — postulates (read terse over)
   * not always correct

Thompson — measured charge-to-mass ratio of electron
   \[ \frac{1.76 \times 10^8 \text{C}}{9} \] * discovered electrons

Millikan — “oil drop experiment”
   * computed electron mass \[ 9.1 \times 10^{-28} \text{g} \]

Thompson — “plum pudding”
   distributed mass of positive charge

Butterfield — “gold foil” experiment
   * disproved plum pudding — atom has very dense positive nucleus with negative electrons (did not know anything about neutrons)

Mass number

\[ ^{12}_6 \text{C} \]

Neutrons = Mass number — atomic number

Atomic number (\# protons = \# electrons)

Isotope \[ ^{12}_6 \text{C} \]

Mass number changes, but atomic number does not change
Molecular mass - molecule; nonmetals/metalloids (amu)
Formula mass - formula unit; ionic (amu)

Average atomic mass

\[
\frac{12 \text{ C (99.49\%)} + 13 \text{ C (1.11\%)}}{(0.9949)(12.00) + (0.0111)(13.00)} = 12.01 \text{ amu}
\]

Molecular compounds
- chemical bonds
- nonmetals
- diatomic molecules

Ionic compounds
- contains both cation and anion
  - cation - positive ion
  - anion - negative ion
- net charge = 0

Periodic table
1a - 1st column/group - alkali metals
2a - 2nd group - alkaline earth metals
7a - halogens
8a - noble gases

*Groups have similar electron arrangements and physical/chemical properties
A) Binary Inorganic Compounds

- Ionic compounds
  - Metal (cation type) + Nonmetal
    \[
    \text{\(\text{Al}^{3+}\), \(\text{Ca}^{2+}\), \(\text{Na}^+\)}
    \]
  - Metal (2+ type) + Nonmetal
    \[
    \text{\(\text{Cu}^2\), \(\text{Cl}^-, \text{Br}^-\)}
    \]

- Different oxidation states:
  - Copper reactions hydrates

B) Molecular compounds

- Nonmetal + Nonmetal
  \[
  \text{\(\text{Fe(OH)}_2\)}
  \]

B) Polyatomic Ions

Know the names:

- Hydroxide \(\text{OH}^-\)
- Ponside \(\text{O}_2^-\)
- Nitrate \(\text{NO}_3^-\)
- Nitric \(\text{NO}_2^-\)
- Carbonate \(\text{CO}_3^{2-}\)
- Phosphate \(\text{PO}_4^{3-}\)
- Perrymagete \(\text{MnO}_4^-\)
- Acetic \(\text{CH}_3\text{COO}^-\)
- Perchlorate \(\text{ClO}_4^-\)
- Chlorate \(\text{ClO}_3^-\)

Cyanide \(\text{CN}^-\)
Thiocyanide \(\text{SCN}^-\)
Sulfate \(\text{SO}_4^{2-}\)
Sulfide \(\text{SO}_3^-\)
Dichromate \(\text{Cr}_2\text{O}_7^{2-}\)
Chromate \(\text{CrO}_4^{2-}\)
Ammonium \(\text{NH}_4^+\)
Mercury (II) \(\text{Hg}_2^{2+}\)
Chloride \(\text{Cl}^-\)
Hypochlorite \(\text{ClO}^-\)
A) Binary Aqueous Acids

\[ H + \text{ion (no oxygen)} \rightarrow \text{H}^+ + \text{ion} \]

Oxo acids

\[ H + \text{polyatomic ion (has oxygen)} \]

\[ H_2SO_4 \quad \text{H}_2PO_4 \quad \text{H}_3PO_4 \quad \text{H}_2SO_3 \]

\[ \text{Nitrate} \rightarrow \text{NO}_3^- \]

\[ \text{Nitrite} \rightarrow \text{NO}_2^- \]

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Chapter 3

* You should know how to balance chemical reactions.

Combustion reaction - 2+ reactions form one product.

Decomposition reaction - one reactant \( \rightarrow \) 2+ products.

* Combustion reaction - Hydrocarbon + \( O_2 \rightarrow \text{CO}_2 + H_2O \)

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Calculate percent composition:

\[ \left( \frac{\text{# atoms of element}}{\text{formula weight}} \right) \times 100 \]

\[ \text{Example: NO}_2 \]

\[ \% \text{Oxygen} = \frac{2 \times (16.00)}{14.01} \times 100 \]

\[ \% \text{Nitrogen} = \frac{1 \times (14.01)}{14.01} \times 100 \]
Mole - quantity = \(6.022 \times 10^{23}\)

Molar mass \(\frac{9}{\text{mole}}\) - given in periodic table

* You should be already comfortable converting mass \(\leftrightarrow\) moles

**Empirical Formula** - simplest integer ratio

* Use percent compositions, and treat as grams

Given:

C 44.77% \(\rightarrow\) 44.77 g

H 7.52% \(\rightarrow\) 7.52 g

O 47.71% \(\rightarrow\) 47.71 g

* Convert to moles of each

\[
\begin{align*}
44.77\text{ g C} & \rightarrow 3.227 \text{ mol C} \\
7.52\text{ g H} & \rightarrow 7.46 \text{ mol H} \\
47.71\text{ g O} & \rightarrow 2.182 \text{ mol O}
\end{align*}
\]

* Now divide by smallest of three

\[
\frac{3.227}{2.182} \quad \frac{3.227}{2.182} \quad \frac{3.227}{2.182}
\]

\[
\Rightarrow \quad \text{C}_5\text{H}_{10}\text{O}_4
\]

\[
\text{C}_6\text{H}_{12}\text{O}_6 - \text{molecular} \\
\text{C}_2\text{H}_4\text{O} - \text{empirical}
\]

Molecular mass of empirical formula = 134.13 g/mol

If you want to determine molecular formula, you need to know the actual total mass of sample/comound, and then multiply empirical formula accordingly.
Limiting Reactant

- Determine mole of reactants, then divide by coefficients
- Smallest value is your limiting reactant

Example: 

\[ \text{CH}_3\text{OH} + 3\text{HBr} \rightarrow \text{CH}_3\text{Br} + \text{H}_2\text{O} \]

5.00 g \( \text{CH}_3\text{OH} \)
10.00 g \( \text{HBr} \)

Determine limiting reactant

\[
\frac{5.00 \text{ g CH}_3\text{OH}}{16.00 \text{ g CH}_3\text{OH}} = 0.3125 \text{ mol CH}_3\text{OH}
\]

\[
\frac{10.00 \text{ g HBr}}{109.9 \text{ g HBr}} = 0.0910 \text{ mol HBr}
\]

Now, if we wanted to determine mass of \( \text{CH}_3\text{Br} \) produced:

\[
0.1236 \text{ mol HBr} \leftrightarrow (0.1236) \text{ mol CH}_3\text{Br}
\]

\[
11.73 \text{ g CH}_3\text{Br}
\]

Be aware of limiting reactants when having to determine percent yield:

\[
\text{percent yield} = \frac{\text{actual}}{\text{theoretical}} \times 100
\]
Chapter 4

Solution = has solvent + solute

- Ions disassociate in water (electrolytes)
- Molecular compounds do not (nonelectrolytes)

Strong acids \(\rightarrow\) strong electrolytes
Weak acids \(\rightarrow\) weak electrolytes

* Know your strong acids; everything else is a weak acid

Monoprotic acid = one \(H^+\) produced per molecule
Diprotic acid = two \(H^+\) produced per molecule

Organic acids = contain carbon

BNINCLS

* Precipitation reactions

Know your solubility rules, and exceptions

Exchange reactions
exchange of atoms in products and reactants

Ionic equations break into ions
Molecular equations \(\rightarrow\) complete ionic equations

Ions + Ag(s) Get rid of spectator ions \(\rightarrow\) net ionic equation
* Ions in gaseous state on both sides

Neutralization reactions

* Acid + Base \(\rightarrow\) \(H_2O\) + ionic salt
Acid + Base \(\rightarrow\) gas + ionic salt
Molarity = \( \frac{\text{Moles solute}}{\text{Volume (L) of solution}} \)

- Know how to use molarity in conversions

**Titrations**

At equivalence point, moles A = moles B
- usually on acid \( \Rightarrow (\text{Molarity of A})(V_A) = (\text{Molarity of B})(V_B) \)
- and a base

**Dilutions**

\[ \frac{M_{\text{concentrated}}}{V_{\text{concentrated}}} = \frac{M_{\text{diluted}}}{V_{\text{diluted}}} \]

Examples:
- \( M_{\text{diluted}} = 0.10 \text{ M} \)
- \( V_{\text{diluted}} = 250 \text{ mL} \)
- \( M_{\text{concentrated}} = 5.0 \text{ M} \)

\[ \Rightarrow V_{\text{concentrated}} = 5 \text{ mL} \]

Thus, 245 mL of water must be added.

**Oxidation-reduction Reactions**

LEO - loses electrons = oxidation

GER - gains electrons = reduction

- Know rules about oxidation states

- On the exams, you may have to identify the type of reaction (combination, decomposition, acid-base, oxidation-reduction)

If there is a change in oxidation numbers, then it is a redox reaction.