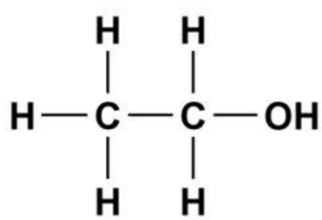
AP Chemistry Unit 1 Notes
Chapters 1 -3

<u>Chapter 1: Matter & Measurement</u>

The Study of Chemistry

- Matter
  - Has mass and occupies space
- Properties
  - Characteristics of matter
  - Allows us to identify and distinguish types of matter
  - Relate to the composition and structure of matter
  - Example: A banana is yellow and soft can change by browning
- Composition
  - Describes what matter is made of
  - Indicates relative proportions
  - Example: Water is H_2O
- Molecules
  - 2 or more atoms joined together by covalent bonds
  - Can be the same element (H_2 & O_2) or different (H_2O)
  - "Minor differences" lead to vastly different properties (ethanol vs. ethylene glycol and water vs. hydrogen peroxide)



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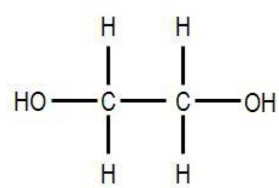
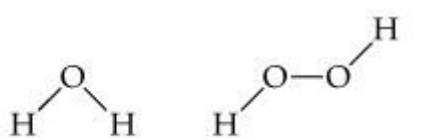
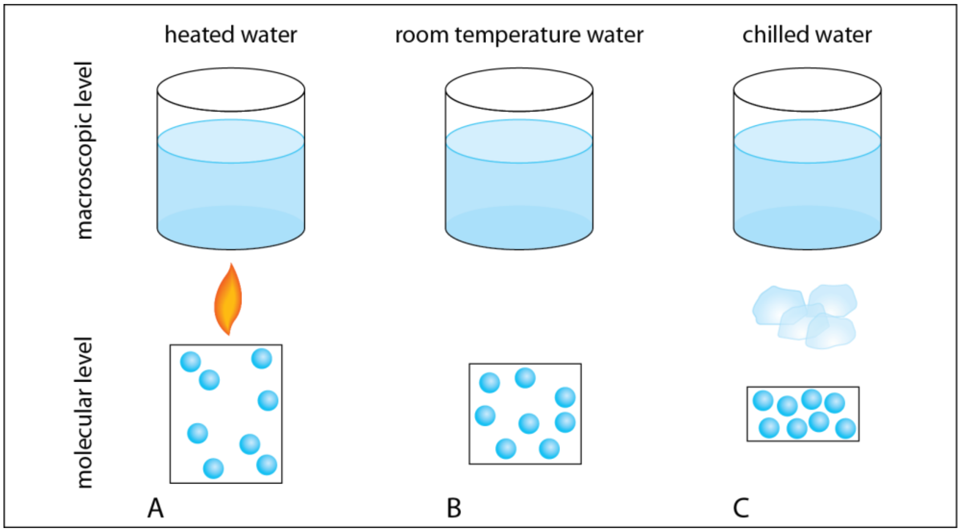


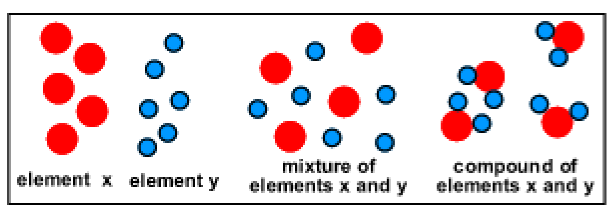


Water Hydrogen peroxide

- Macroscopic
  - Refers to ordinary sized objects
  - This is where we make our observations
- Submicroscopic
  - Refers to things on the scale of atoms & molecules
  - Changes in structure at this level cause the changes that we observe at the macroscopic level
  - Particulate diagrams help to illustrate/model this level



- Classification of Matter
  - Solid:
    - Definite shape & volume
    - Particles packed tightly
    - Vibrate in fixed position
  - Liquids
    - Definite volume
    - Takes shape of container
    - Particles allowed to move around each other (flow)
  - Gas
    - Take the shape & volume of container
    - Particles move freely
    - Can be compressed and expanded
- Pure Substances
  - Have distinct properties
  - Composition does not vary from sample to sample
  - Includes
    - Elements (only one type of atom)
    - Compounds (chemical combination of 2 or more elements)
- Mixtures
  - Combinations of 2 or more substances
  - Properties of components are not changed
  - Compositions vary between samples



- Types of Mixtures
  - Heterogeneous
    - Material varies throughout
    - Distinct visible layers/parts
  - Homogeneous
    - Material is uniform throughout

        - Also called solutions
        - Can be solid, liquid, or gas
- Law of Constant Composition
    - The elemental composition of a pure compound is always the same
- Properties of Matter
    - Physical Properties
        - Determined without changing the identity of the substance
        - Can be measured or described with the senses
    - Chemical Properties
        - Describe how substances can react to form new substances
    - Intensive Properties
        - Do not depend on the amount of substance
        - Ex. boiling point, density, color
        - Can be used to identify substances
    - Extensive Properties
        - Depend on the amount (quantity) of substance
        - Ex: Mass, volume, length
- Physical Changes
    - Substance changes its appearance but not its composition
    - All Phase changes are Physical
- Chemical Changes
    - Substances is transformed into a different substance
    - Chemical reactions
    - Copper + Nitric Acid (Story)
- Separation of Mixtures
    - Filtration
    - Distillation
    - Chromatography
    - Decantation
- Units of Measurement
    - SI Base Units (Table 1.4 page 14)

Dimension	Unit	Abbreviation
Mass	kilogram	kg
Length	meter	m
Time	second	s
Temperature	Kelvin	K
Amount of substance	Kelvin	mol
Electric current	Ampere	A
Luminous intensity	Candela	cd

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- Selected Metric Prefixes (Table 1.5 page 14)

Prefix	Abbreviation	Meaning
giga	G	10^9
mega	M	10^6
kilo	k	10^3
deci	d	10^{-1}
centi	c	10^{-2}
milli	m	10^{-3}
micro	μ	10^{-6}
nano	n	10^{-9}
pico	p	10^{-12}
femto	f	10^{-15}

- Temperature Conversions
    - $K = {}^oC + 273$
    - ${}^oC = 5/9({}^oF – 32)$
    - ${}^oF = 9/5{}^oC + 32$

SAMPLE EXERCISE 1.3 Page 16

If a weather forecaster predicts that the temperature for the day will reach 31^{o}C, what is the predicted temperature in (a) K and (b) ^{o}F?

PRACTICE 1.3 Page 16

- Derived Units
  - Volume
    - cm^3
    - mL
  - Density
    - g/cm^3
  - Speed
    - m/s

SAMPLE EXERCISE 1.4 Page 20

(a) Calculate the density of mercury if 1.00×10^2 g occupies a volume of 7.36 cm^3.

(b) Calculate the volume of 65.0 g of the liquid methanol (wood alcohol) if its density is 0.791 g/mL.

(c) What is the mass in grams of a cube of gold (density = 19.32 g/cm^3) if the length of the cube is 2.00 cm?

PRACTICE EXERCISE 1.4 Page 20

- Uncertainty in Measurement
- Precision vs. Accuracy
    - Precision
        - Degree of closeness of several measurements to each other
    - Accuracy
        - Degree of closeness of a single measurement to the "true" or accepted value
- Significant Figures
    - Measured quantities are reported in a way that only the last digit is uncertain

- Which Digits are Significant?

Digits	When to Count	Example
Nonzero	Always	2.514 = 4 sig figs
Leading zeroes	Never	0.025 = 2 sig figs
Trailing zeroes	After a decimal	250 = 2 sig figs 2.50 = 3 sig figs
Captive zeroes	Always	200.5 = 4 sig figs 200.0 = 4 sig figs

SAMPLE EXERCISE 1.6 Page 23

How many significant figures are in each of the following numbers (assume that each number is a measured quantity)?

(a) 4.003
(b) 6.02×10^{23}
(c) 5000

PRACTICE EXERCISE 1.6 Page 23

- Calculating with Sig Figs:
  - Adding or Subtracting
    - Report answer with the same number of decimal places as the least precise number in the calculation
  - Multiplying or Dividing
    - Report answer with the same number of significant digits as the least precise number in the calculation

SAMPLE EXERCISE 1.7 Page 24

The width, length, and height of a small box are 15.5 cm, 27.3 cm, and 5.4 cm, respectively. Calculate the volume of the box, using the correct number of significant figures in your answer.

PRACTICE EXERCISE 1.7 Page 24

SAMPLE EXERCISE 1.8 Page 25

A gas at 25^{o}C fills a container whose volume is 1.05 x 10^{3} cm^{3}. The container plus the gas have a mass of 837.6 g. The container, when emptied of all gas, has a mass of 835.2 g. What is the density of the gas at 25^{o}C?

PRACTICE EXERCISE 1.8 Page 25

- Dimensional Analysis
    - A method used to solve numerical problems and check solutions for possible errors
    - Units are important!
    - Often uses conversion factors:
        - Fractions that have the same quantity with different units in the numerator and denominator (ex. 12 in/1ft or 1 mole/23.0 g Na)

SAMPLE EXERCISE 1.9 Page 26

If a woman has a mass of 115 lbs, what is her mass in grams?

PRACTICE EXERCISE 1.9 Page 26

SAMPLE EXERCISE 1.10 Page 27

The average speed of a nitrogen molecule in air at 25^{o}C is 515m/s. Convert this speed to miles per hour.

PRACTICE EXERCISE 1.10 Page 27

SAMPLE EXERCISE 1.11 Page 28

Earth's oceans contain approximately 1.36 x 10^9 km^3 of water. Calculate the volume in liters.

PRACTICE EXERCISE 1.11 Page 28

SAMPLE EXERCISE 1.12 Page 29

What is the mass in grams of 1.00 gal of water? The density of water is 1.00 g/mL.

PRACTICE EXERCISE 1.12 Page 29

Chapter 2 Atoms, Molecules & Ions

The Atomic Theory of Matter

- Democritus (460 – 370 BC)
  - Greek philosopher
  - Thought that material was made up of tiny particles called atomos
  - Aristotle and Plato formulated idea that there can be no ultimately indivisible particles of matter
  - Atomic view faded
- John Dalton
  - English school teacher
  - Studied meteorology
  - Made observations of atmospheric gases that led to the development of his atomic theory during the period of 1803 – 1807
  - Atomic Theory
    - Elements are composed of extremely small particles called atoms
    - All atoms of a given element are identical to each other. Atoms of one element are different from the atoms of all other elements.
    - Atoms of one element cannot be changed into other elements by chemical reactions. Atoms are not created or destroyed in chemical reactions.
    - Compounds are formed when the atoms of one element combine with the atoms of another element. A given compound always has the same relative number and kind of atoms.
    - Atoms are the smallest part of an element that have the properties of the element

- Law of constant composition: In a compound the relative numbers and types of atoms are constant
- Law of conservation of mass: In a chemical reaction, the masses of the products must be equal to the masses of the reactants
- Law of multiple proportions: If two elements A and B form more than one compound, the masses of B that can combine with A are in the ratio of small whole numbers
    - For example: Water (H_2O) and Hydrogen peroxide (H_2O_2)

      In water 8.0 g of oxygen combines with 1.0 g of hydrogen

      In hydrogen peroxide 16.0 g of oxygen combines with 2.0 grams of hydrogen

Discovery of Atomic Structure

- We now know that atoms are composed of subatomic particles
- The three main subatomic particles are protons, neutrons, and electrons.
- Particles with the same charge repel each other while particles with opposite charges attract one another

JJ Thomson (late 1800's)

- Experimented with a cathode ray tube
- Concluded that cathode rays are streams of negatively charged particles
- Credited with the discovery of the electron
- Calculated charge to mass ratio of the electron (1.76×10^{8} C/g)
- Developed "Plum Pudding Model" of the atom

Roger Millikan (1868 – 1953)

- Oil-drop experiment
- Was able to deduce the charge of the electron (1.602×10^{-19} C)
- Used Thomson's charge to mass ratio to calculate the mass of the electron (9.10×10^{-28} g)

Henri Becquerel (1852 – 1908)

- Discovered radiation from a uranium compound

- Marie and Pierre Curie experimented to isolate the radioactive components of the compound
- Further studies revealed three types of radiation : alpha, beta, gamma

Ernest Rutherford

- Showed that alpha and beta radiation consists of fast moving particles
- Beta particles are high speed electrons (radioactive equivalent of cathode ray)
- Alpha particle have a charge of 2+ and Beta have a charge of 1-
- Gamma radiation is high energy radiation similar to x rays. It does not consist of particles and does not carry a charge.
- Performed gold foil experiment:
    - Atom is mostly empty space
    - Atom contains small, dense, positively charged center called the nucleus

Subatomic Particles

- Proton has a charge of 1.602×10^{-19} C
- Electron has a charge of -1.602×10^{-19} C
- 1.602×10^{-19} C is called the electronic charge
    - We express charges of atomic and subatomic particles in multiples of this number….proton is +1 and electron is -1
    - Every atom has equal numbers of protons and electrons and therefore has no electric charge

Masses of Atoms

- Atoms have very small masses and expressing the mass in grams is not useful, so we use atomic mass units (amu)
- 1 amu = 1.66054×10^{-24} grams
- Proton mass = 1.0073 amu
- Neutron mass = 1.0087 amu
- Electron mass = 5.486×10^{-4} amu

Sizes of Atoms

- Atoms are very small (most have diameters between 1 x 10^{-10} and 5 x 10^{-10} meters)
- Angstroms are often used to measure atomic sizes (1 angstrom = 10^{-10} meters)
- The nucleus is very small compared to the rest of the atom (on the order of 10^{-4} angstroms)

SAMPLE EXERCISE 2.1 Page 44

The diameter of a US penny is 19 mm. The diameter of a silver atom, by comparison, is only 2.88 angstroms. How many silver atoms could be arranged side by side in a straight line across the diameter of a penny?

PRACTICE EXERCISE 2.1 Page 44

Atomic Numbers

- Number of protons in an atom
- Unique for every element
- Because the atom has no electric charge, this is also the number of electrons in a neutral atom

Mass Number

- Refers to the total number of protons and neutrons in an atom

SAMPLE EXERCISE 2.2

How many protons, neutrons, and electrons are in (a) an atom of ^{197}Au (b) an atom of strontium-90?

PRACTICE EXERCISE 2.2

SAMPLE EXERCISE 2.3

Magnesium has three isotopes, with mass numbers of 24, 25, and 26. (a) Write the complete chemical symbol for each of them (b) How many neutrons are in an atom of each isotope?

PRACTICE EXERCISE 2.3

Isotopes

- Atoms of the same element with different numbers of neutrons
- Will also have different masses

- Same number of protons (same element)
- Carbon 12 vs. carbon 14

Determining the Mass of an Element

- Mass Spectrometry (Video)

Average Atomic Masses

- Most elements exist as mixtures of isotopes
- Average atomic mass is determined using masses of various isotopes and their relative abundances (mass spectrograph)
- Average atomic mass is also called atomic weight

SAMPLE EXERCISE 2.4

Naturally occurring chlorine is 75.78% ^{35}Cl, which has an atomic mass of 34.969 amu, and 24.22% ^{37}Cl which has a mass of 36.966 amu. Calculate the average atomic mass (that is, the atomic weight) of chlorine.

PRACTICE EXERCISE 2.4

The Periodic Table

- Tool for chemists to organize and remember chemical facts
- Elements are arranged by increasing atomic number with elements that have similar properties placed in vertical columns (called groups or families)
- Horizontal rows are called periods
- Names of some groups
    - Group 1 (1A) → Alkali Metals
    - Group 2 (2A) → Alkaline Earth Metals
    - Group 16 (6A) → Chalcogens
    - Group 17 (7A) → Halogens
    - Group 18 (8A) → Noble Gases
- Metals are on the left
- Nonmetals are on the right (except hydrogen)
- Metalloids lie along the line and exhibit properties that are between those of metals and nonmetals

SAMPLE EXERCISE 2.5

Which two of the following elements would you expect to show the greatest similarity in chemical and physical properties? B, Ca, He, Mg, P?

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PRACTICE EXERCISE 2.5

Molecules

- A molecule is an assembly of two or more atoms tightly bound together
- Chemical formulas show the number and types of atoms present in a molecule
- Diatomic molecules are made up of two atoms
- Elements that exist as diatomic molecules
    - H
    - N
    - O

- F
- Cl
- Br
- I

- Molecular Compounds
  - Compounds composed of molecules
  - Contain more than one type of atom (different elements)
  - Most molecular compounds will contain only nonmetal atoms
- Molecular & Empirical Formulas
  - Formulas that indicate the actual numbers and types of atoms in a compound are molecular formulas
  - Formulas that give only the relative number of atoms of each type are called empirical formulas
  - For many substances the molecular and empirical formulas are the same

SAMPLE EXERCISE 2.6

Write the empirical formulas for the following molecules: (a) glucose, a substance also known as either blood sugar or dextrose, whose molecular formula is $C_6H_{12}O_6$ (b) nitrous oxide, a substance used as an anesthetic and commonly called laughing gas, whose molecular formula is N_2O.

PRACTICE EXERCISE 2.6

Picturing Molecules (see page 54)

- Structural formula
  - Shows which atoms are attached to which within the molecule
- Perspective drawing
  - Gives some sense of three dimensional shape of the molecule. Solid lines represent bonds in the plane of the paper, the solid wedge is a bond that extends out of the paper and the dotted line is a bond behind the paper.
- Ball and Stick model

    - Atoms as spheres and bonds as sticks. Shows the angles of the bonds.
- Space filling models:
    - Show what the molecules would look like if the atoms were scaled up in size

Ions

- Atom that has a charge due to loss or gain of electrons
- Cation
    - Positively charged
    - Results from loss of electron
- Anion
    - Negatively charged
    - Results from gain of electron
- In general metal atoms will lose electrons and nonmetal atoms will gain electrons
- Polyatomic Ions
    - Atoms bonded as in a molecule but with an overall electric charge
- Predicting charges of ions:
    - Group 1 ➔ 1+
    - Group 2 ➔ 2+
    - Group 17 ➔ 1-
    - Group 16 ➔ 2-
    - Nitrogen ➔ 3-
    - Aluminum ➔ 3+

SAMPL E EXERCISE 2.7

Give the chemical symbol, including mass number, for each of the following ions: (a) the ion with 22 protons, 26 neutrons, and 19 electrons (b) the ion of sulfur that has 16 neutrons and 18 electrons

PRACTICE EXERCISE 2.7

SAMPLE EXERCISE 2.8

Predict the charge expected for the most stable ion of barium and for the most stable ion of oxygen.

PRACTICE EXERCISE 2.8

Ionic Compounds

- Contain both positively and negatively charged ions
- Generally combinations of metals and nonmetals
- Combine so that overall electric charge is neutral

SAMPLE EXERCISE 2.9

Which of the following compounds would you expect to be ionic? N_2O, Na_2O, $CaCl_2$, SF_4?

PRACTICE EXERCISE 2.9

SAMPLE EXERCISE 2.10

What are the empirical formulas of compounds formed by (a) Al^{3+} & Cl^- (b) Al^{3+} & O^{2-} (c) Mg^{2+} & NO_3^-

PRACTICE EXERCISE 2.10

Naming Cations

- Cations formed from metals have the same name as the metal
- If a metal can have more than one ionic charge, the charge is indicated by a Roman numeral in parentheses following the name of the metal
- Sometimes the –ous or –ic suffix is added to the Latin name to indicate charge
- Cations formed from nonmetal will end with -ium suffix

Naming Anions

- The names of monatomic anions are formed by replacing the ending of the name of the element with –ide
- Polyatomic ions containing oxygen (oxyanions) have names ending with –ate or –ite
- Anions derived by adding H^+ to an oxyanion are named by adding as a prefix the word hydrogen or dihydrogen as appropriate

Ionic Compounds

- Named by putting the name of the cation with the name of the anion

Acids

- Acids are hydrogen containing compounds
- An acid will consist of an anion with enough hydrogen ions to balance or neutralize it
- Acids with anions whose name ends with –ide are named by changing the –ide to –ic, and adding the prefix hydro- with the word acid
- Acids with anions whose names end with –ate or –ite are named by changing the ending t0 –ic or –ous and adding the word acid

Binary Molecular Compounds

- The name of the element farther to the left on the periodic table is usually written first (exception is oxygen which is always written last except when bonded with fluorine)
- If both are in the same group, the one with the higher atomic number is named first
- The name of the second element is given an –ide ending
- Greek prefixes are used to indicate the numbers of atoms (except mono- is not used for the first element)
    - 1 ➔ mono-
    - 2 ➔ di-
    - 3 ➔ tri-
    - 4 ➔ tetra-
    - 5 ➔ penta-
    - 6 ➔ hexa-
    - 7 ➔ hepta-

- 8 ➔ octa-
- 9 ➔ nona-
- 10 ➔ deca-

SAMPLE EXERCISE 2.12

Name the following compounds: (a) K_2SO_4 (b) $Ba(OH)_2$ (c) $FeCl_3$

PRACTICE EXERCISE 2.12

SAMPLE EXERCISE 2.13

Write the chemical formulas for the following compounds: (a) potassium sulfide (b) calcium hydrogen carbonate (c) nickel (II) perchlorate

PRACTICE EXERCISE 2.13

SAMPLE EXERCISE 2.14

Name the following acids: (a) HCN (b) HNO_3 (c) H_2SO_4 (d) H_2SO_3

PRACTICE EXERCISE 2.14

SAMPLE EXERCISE 2.15

Name the following compounds: (a) SO_2 (b) PCl_5 (c) N_2O_3

PRACTICE EXERCISE 2.15

- Metals are on the left (except hydrogen)
- Nonmetal are on the right
- Metalloids lie along the line and exhibit properties that are between those of metals and nonmetals