# ***Chemistry***

## *1st Semester Exam Review Notes*

**Chapter 1**

**1.1 - major areas of study**
- analytical, physical, biochemistry, organic, inorganic

**1.3 - scientific method**
- observation, hypothesis, experiment, theory/conclusion

**Nature & philosophy of science**
- inductive vs. deductive thinking
  - inductive – goes from observable facts to theory
  - deductive – goes from theory to testable statements
- abductive reasoning – inference to the best explanation
- falsificationism – falsification, falsifiability
  - the best scientific theories are easily falsifiable but have not yet been falsified
- paradigms – accepted sets of rules used by scientists
  - paradigm shifts – abrupt changes in accepted set of rules
  - prescience → normal science → revolutionary science
  - the purpose of science is to solve puzzles; the theories that solve the most puzzles with the fewest anomalies are considered to be 'the best'

**Chapter 2 - Matter & Change**

**2.1 - physical properties, states of matter, physical changes**
- physical properties - can be observed without permanent change
- states of matter - solid, liquid, gas (in order of increasing energy)
- physical changes - non-permanent change in a substance/matter

**2.2 - heterogeneous vs. homogeneous mixtures, separating mixtures**
- heterogeneous mixtures - not uniform in composition (i.e., salad)
- homogeneous mixtures - uniform in composition; solutions (i.e., Kool-aid); alloys (mixtures of metals)
- separating mixtures - distillation/boiling, magnetism, sifting, filtration, dissolving

**2.3 - elements vs. compounds, chemical symbols**
- elements - simplest form of matter
- compounds - chemical combinations of elements
- chemical symbols – always have a first capital letter, usually a second lowercase letter

**2.4 - reactants vs. products, chemical properties, law of conservation of mass**
- reactant(s) – what your reaction starts with; left side of chemical equation
- product(s) – what your reaction makes; right side of chemical equation
- chemical properties – can only be observed through chemical reaction
- law of conservation of mass
  - mass must be conserved in a chemical reaction
  - mass of reactants = mass of products
  - matter cannot be created nor destroyed, only changed

## Chapter 3 - Scientific Measurement

**3.1 - qualitative vs. quantitative measurements, scientific notation**

- qualitative – non-numeric measurement
- quantitative – numeric measurement
- scientific notation – very small and very large numbers represented with powers of 10
  - $6.02 \times 10^{23}$ = 602,000,000,000,000,000,000,000
  - $4.7 \times 10^{-12}$ = 0.000 000 000 004 7

**3.2 - accuracy, precision, error**

- accuracy – how close your measurement is to a scientifically accepted value
- precision – how close your measurements are to one another
- error – how far off your measurement is from the accepted value
- $\boxed{\%\text{ error} = \left(\frac{|\text{accepted} - \text{experimental}|}{\text{accepted}}\right) * 100\%}$ (p. 77)

**3.3 - SI (metric system), volume, mass**

- SI – system of measurement based on powers of 10; prefixes
  - meter, kilogram, second, $^{O}$C (or K)
- volume - derived unit from lengths of 3 dimensions; $\boxed{V = L \times W \times H}$
- mass – amount of matter in your sample; measured in grams

**3.4 – density**

- density – how much matter is in a given volume; $\boxed{D = \frac{m}{V}}$

**3.5 - temperature (°C vs. K)**

- 0 °C – freezing point of water
- 100 °C – boiling point of water
- 37 °C – human body temp.
- Kelvins – absolute zero scale; no negative temps.; K > °C
  - $\boxed{K = °C + 273}$
  - $\boxed{°C = K - 273}$

## Chapter 4 - Problem Solving in Chemistry

**4.2 - conversion factors, metric conversions** – goalpost method

**4.3 - multistep conversions** – add another goalpost

## Chapter 5 - Atomic Structure and the Periodic Table

**5.1 - Dalton's atomic theory, size of an atom**

- four parts to Dalton's atomic theory; relative size of an atom (very small!)

**5.2 - electrons, cathode ray tube, Thomson, protons, neutrons, nucleus, Rutherford's gold foil experiment**

- Thomson – discovered the electron; cathode ray tube experiment; electron is very very tiny and has a negative charge; "plum pudding model"
- Rutherford – gold foil experiment; credited with discovering the nucleus; nucleus is the tiny center of the atom with all the mass; positive, made of protons (+) and neutrons (0)

**5.3 - atomic number, mass number, isotopes, atomic mass**

- atomic number = # of protons; each element has a unique atomic number
- mass number = # of protons + # of neutrons; represents the number of particles in the nucleus
- isotopes – elements with more or less neutrons than usual
  - carbon-12 has a mass number of 12; 6 $p^+$, 6 $n^0$
  - carbon-14 has a mass number of 14; 6 $p^+$, 8 $n^0$
- atomic mass – average of isotope masses for a given element based on prevalence

**5.4 - periodic table, Mendeleev, periods, periodic law, groups, representative elements, metals vs. nonmetals, alkali metals, alkaline earth metals, transition metals, inner transition metals, halogens, noble gases, metalloids**

- periods – rows; same outer energy level = row #
- groups – columns; same # of valence electrons; share similar chemical properties
- periodic law (aka periodicity) – trends emerge when the elements are arranged by group & period
- metals vs. nonmetals; metalloids; representative (A) vs. transition (B) elements
- Group 1 (IA) – alkali metals
- Group 2 (IIA) – alkaline earth metals
- Group 17 (VIIA) – halogens
- Group 18 (VIIIA or 0) – noble gases
- Group B – transition metals
- Lanthanide & Actinide series – inner transition metals

## Chapter 13 - Electrons in Atoms

**13.1 - atomic models (Dalton, Thomson, Rutherford, Bohr, quantum mechanical), atomic orbitals**

- Bohr – electrons orbit the nucleus in distinct paths (like planets); atomic spectra
- Quantum mechanical – electrons exist somewhere in orbitals; currently accepted model
- orbital – area around the nucleus where an electron is likely to be found

**13.3 - line spectrum, atomic emission spectra**

- line spectrum – caused by release of light by excited electrons jumping down energy levels
- atomic emission spectrum – allows identification of elements

## Chapter 6 - Chemical Names and Formulas

**6.1 - molecules/molecular compounds vs. ions/ionic compounds, cation vs. anion, naming ions**

- molecules – covalent compounds; electrons are shared; weak bonds; individual units
- ionic compounds – electrons are transferred; strong bonds; can be dissolved by water; crystal lattice w/formula units

**6.2 - chemical formulas, molecule vs. formula unit, laws of definite and multiple proportions**

- law of definite proportions – masses of elements are always in the same proportions (mass ratio) in a given compound
- law of multiple proportions – elements can combine in different mass ratios to form different elements

**6.3 - monatomic vs. polyatomic ions, ionic charges, transition metal ions**

- monatomic – only one atom
- polyatomic – more than one atom
- metals have positive charges; nonmetals have negative charges; polyatomic ions tend to have negative charges
- transition metals – can have more than one charge; shown with Roman numeral

**6.4 - writing formulas & names for ionic compounds (criss-cross method)**

- for formulas to names: write name of positive ion then negative ion
- for names to formulas: write symbols of positive & negative ions with charges; criss-cross charges if they don't cancel out; put parentheses around multiple polyatomic ions

**6.5 - writing formulas & names for molecular compounds (prefixes), common acids**

- for formulas to names: write first element name w/prefix, then second ionic name w/prefix
- for names to formulas: write element symbol w/subscripts for each element

## Chapter 7 - Chemical Quantities

**7.1 - mole, formula weight/molar mass**

- 1 mole = $6.02 \times 10^{23}$ particles
- molar mass = sum of molar masses of all elements in a compound, measured in g/mol

**7.2 - molar mass, molar volume, converting between moles and other units**

- molar volume = 22.4 L/mol at Standard Temperature and Pressure (STP)

- Figure 7.13

**7.3 - percent composition**

- $$\boxed{percent\ composition = \frac{grams\ of\ element}{total\ grams\ of\ compound} x100\%}$$

## Chapter 8 - Chemical Reactions

**8.1 - word equations, chemical equations, catalyst, balancing chemical equations, coefficients**

- reactants $\rightarrow$ products
- catalyst – speed up reaction but are not used in reaction
- additional symbols – *(s)*, *(l)*, *(g)*, *(aq)*, $\Delta$

**8.2 - types of chemical reactions**

- synthesis/combination – two reactants make one product
- decomposition – one reactant makes 2+ products
- single replacement – an element reacts with an ionic compound
- double replacement – two ionic compounds swap positive ions
- combustion – a chemical reacts with oxygen gas

## Chapter 9 - Stoichiometry

**9.1 - meaning of coefficients in balanced equations**

- coefficients – represent relative number of moles of reactants and products

**9.2 - mole ratio, conversions between moles and other units**

- mole ratio – convert from moles of ‘given’ to moles of ‘wanted’
- convert from grams to moles by dividing by molar mass
- convert from moles to grams by multiplying by molar mass

**9.3 - limiting reagent, percent yield, theoretical yield, actual yield**

- limiting reagent – runs out first, determines the amount of product formed
- percent yield – ratio of actual yield to theoretical yield, percentage of expected product formed
- theoretical yield – maximum amount of product expected, determined by stoichiometry
- actual yield – amount of product formed in laboratory