CHM 150/151: Final Exam Study Guide

Chapter 1: Matter and Measurement

- Scientific method
- Significant Figures (Sig Figs)
- Scientific notation
- Metric system
  - Know prefixes from nano to Giga
  - Carry out metric-metric conversions
  - 1 $cm^3 \equiv 1$ mL and 1 $dm^3 \equiv 1$ L
- Know 1 in. = 2.54 cm
- Volume by displacement
- density = $\frac{\text{mass}}{\text{volume}}$
  - Be able to determine density, mass, or volume given the other two quantities
- Temperature
  - Perform conversions for °F-to-°C or °C-to-°F and K-to-°C or °C-to-K
- Length, mass, weight, volume
- Know **law of conservation of mass** and solve corresponding problems.
- Classification of matter
  - Given examples or molecular-level images, identify elements, compounds, and mixtures, as well as solids, liquids, and gases.
- Identify properties and changes as physical or chemical
- Solve problems using **unit analysis method**

Chapter 2: Atoms, Molecules, and Ions

- Subatomic particles
  - proton (p^+): +1 charge, inside nucleus, ~1 amu
  - neutron (n): neutral, inside nucleus, ~1 amu
  - electron (e^-): –1 charge, outside nucleus, 0 amu
- Mass number, atomic number, element symbol
  - Given 2, be able to determine the missing info
- Recognize definition of the term **isotope**
- Atomic Notation (or Nuclear Symbol):

  $$\text{mass number} = A,\ \text{atomic number} = Z:\quad {}^{A}_{Z}E \text{ = element symbol}$$
  - Be able to give nuclide symbol for any element given element name and mass number
- Determine the # of protons, neutrons, and electrons given a specific isotope
- Periodic Table:
  - Know terms: groups, periods, Main-group (Representative) elements, transition metals
  - names of Groups IA, IIA, VIIA, and VIIIA

Chapter 2 (Continued)

- ***Know the names and symbols of the first 20 elements of the Periodic Table plus*** *Cr, Mn, Fe, Co, Ni, Cu, Zn^{+2}, Br, Kr, Sr, Ag^{+1}, Cd^{+2}, I, Ba, Hg, Au, Sn, Pb*
- Metals, nonmetals, and semimetals:
  - Location on Periodic Table and properties
- Know which elements are solids, liquids, gases at room temperature (25°C)
- Know which elements exist as **diatomic molecules** at 25°C (H_2, N_2, O_2, F_2, Cl_2, Br_2, I_2)
- **molecule**: compound of *2 or more nonmetals*
  - formula indicates # of each atom present
- **ionic compound**: consists of *metal + nonmetal(s)*
  - *formula unit* indicates ratio of ions present
- **Know ions formed by Main-Group elements**
  - Metals lose electrons → cations
  - Nonmetals gain electrons → anions
  - Group I: +1 charge as ion
  - Group II: +2 charge as ion
  - Group V: -3 charge as ion
  - Group VI: -2 charge as ion
  - Group VII: -1 charge as ion
- **KNOW names and formulas for most common POLYATOMIC IONS**
- **Nomenclature:**
  - Identify a compound as ionic or molecular
  - Given formula of a compound, determine name.
  - Given name of a compound, determine formula.
- **Molecular Compounds:** all nonmetals
  - Use Greek prefixes
  - Know water, ammonia, methane, hydrogen peroxide
- **Acids:** Have H in front, physical state is (aq)
  - Be able to name acids given the formula or determine the formula given the name
- Be able to distinguish between ionic compounds, acids, bases
  - Know why NH_3 is considered a base.

Chapter 3: Formulas, Equations, and Moles

Molar Masses

– Calculate molar mass of elements & compounds

1 mole = 6.022×10^{23} units (Avogadro's #) = molar mass of substance

Chapter 3 (Continued)

Standard Temperature and Pressure (STP):
– Temperature=0°C and Pressure=1 atm

Molar Volume: 1 mole of **gas** occupies **22.4L** at **STP**

Mole Conversions
– Convert among masses, molar mass, moles, molar volume, molecules, atoms

Percent Composition (or Mass Percentage)
– Calculate the percent by mass of each element in a compound given the formula

Empirical and Molecular Formulas
– Determine given masses of elements: using law of conservation of mass or for the mass percentage data and molecular weight, determine the empirical formula and molecular formula

Balancing Chemical Equations
– For various reactions including combustion reactions with odd # of oxygen molecules reacting

Stoichiometry: Calculations involving amounts of reactants and products in a chemical reaction
– coefficients give mole ratios
– Mole-mole, mass-mass, mass-volume, or volume-volume calculations

Limiting Reagent Problems
– Calculate the mass or volume of product that can be made from each of given amount of reactants and using chemical equation
– Smallest amount of product = amount produced
–Reactant producing smallest = limiting reagent
–All other reactants = reactants in excess
– Calculate the mass or volume of reactant in excess that remains after the reaction

Molarity, Solution Stoichiometry, and Volumetric Analysis Calculations
– Solve for molarity given solute present
– Use molarity and volume to solve for moles
– Recognize the concentration of ions given original concentration of ionic solid that dissociates
– Use chemical equation given to determine mole-to-mole ratios needed to solve problems
– Solve for molarity, molar mass, or volume needed to reach equivalence point when given titration data

Mass Percent Concentration

$$\text{M/M\%} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100\%$$

– Solve for or use mass percent concentration in various calculations

Yields of Reactions
<u>theoretical yield</u>: amount of product predicted by balanced equation when limiting reagent is used up
<u>actual yield</u>: amount of product one actually gets

$$\text{Percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

Chapter 4: Reactions in Aqueous Solutions

- **Aqueous Solutions:** ions or compounds (solute) dissolved in H_2O (solvent)
- <u>**Classify and balance Chemical Reactions**</u>
  - **Precipitation Reactions**
    - **precipitate (ppt):** solid forming when 2 ionic solutions react
  - **Acid-Base Neutralization Reactions**
    - Know Arrhenius and Bronsted-Lowry (B-L) definitions for acids and bases
    - Identify B-L acid/base in reaction
  - **Oxidation-Reduction (Redox) Reactions**
    - **Combination reactions**
    - **Decomposition reactions**
    - **Single-replacement reactions**
      - solid metal + metal solution
      - solid metal + acid
      - solid metal + H_2O (l)
    - **Combustion reactions**
    - $C_xH_y + O_2 \rightarrow CO_2\,(g) + H_2O\,(g)$
    - $C_xH_yO_z + O_2 \rightarrow CO_2\,(g) + H_2O\,(g)$

• Given an **Oxidation-Reduction (Redox) Reaction,** use oxidation numbers to identify what reactant is oxidized (lost electrons=LEO, acts as reducing agent) and which reactant is reduced (gains electrons=GER, acts as oxidizing agent)

- **Predict products given set of reactants for:**
  - **Precipitation Reactions** (use Solubility Rules)
  - **Acid-Base Neutralization Reactions**
    - HX + MOH → H_2O (l) + salt
- **Predict products given a set of reactants for Oxidation-Reduction (Redox) Reactions**
  - **Combination reactions**
  - **Decomposition reactions**
  - **Single-replacement reactions**
    - solid metal + metal solution (Activity Series)
    - solid metal + acid (Activity Series)
    - solid metal + H_2O (l) (Active Metals)
  - **Combustion reaction**
  - $C_xH_y + O_2 \rightarrow CO_2\ (g) + H_2O\ (g)$
  - $C_xH_yO_z + O_2 \rightarrow CO_2\ (g) + H_2O\ (g)$
- **Electrolyte:** substance producing ions when dissolved in water to conduct electricity
  - **strong electrolyte:** breaks up completely
    - → many ions present to conduct electricity
    - e.g. strong acids & bases, aqueous salts
  - **weak electrolyte:** breaks up to small degree
    - → only few ions present to conduct electricity
    - e.g. weak acids & bases, insoluble salts
  - **nonelectrolyte:** a molecular compound that forms molecules in water
    - → no ions → does not conduct electricity
  - Be able to classify a substance as a strong, weak, or nonelectrolyte
  - Know strong acids and bases, all others weak

Chemical Equations and **(Net) Ionic Equations**
- **Chemical Equation:** compounds shown as whole
- **Complete/Total Ionic Equation:**
  - shows strong electrolytes as separate ions
- **Spectator Ions:** unchanged in chemical rxn
- **Net Ionic Equation:** Shows what substances change in a chemical reaction (no spectator ions are present)
- Be able to write net ionic equations

Chapter 5: Periodicity and Atomic Structure

wavelength (λ): distance between successive peaks; generally given in m, cm, or nm
- Know 1 m = 10^9 nm

frequency (ν): number of waves passing by a given point in 1 s; given in $\frac{1}{s}$=hertz (Hz)

speed of light, c: 3.00×10^8 m/s

Use equations: $c = \lambda\nu$ and $E = h\nu = \frac{hc}{\lambda}$ to relate energy, frequency, and wavelength

Dual Nature of the Electron
- Electron can behave like a wave or a particle

Write **electron configurations** for elements **1-56**.
- Be able to use the periodic table to know order of orbitals in terms of filling
- Be able to write using full notation and core notation (noble gas or short-hand notation)
- Recognize extra stability gained with filled and half-filled d orbitals

Bohr Model of the Atom
- Electrons move in quantized orbits called "energy levels" around nucleus
- ground state: electron in lowest energy orbitals
- excited state: electron in higher energy orbitals
- When atom absorbs energy, electron jumps from a lower energy to a higher energy orbit.
- When an atom drops from a higher energy to a lower energy orbit, it releases energy, as light in some cases → atomic emission spectra

Atomic Orbital Shapes:
- reflect the "probability density" for an electron in a given orbital
- Know general **shapes for s, p,** and **d orbitals**

Be able to draw **atomic orbital diagrams**
- Know Pauli Exclusion Principle
- Know Hund's Rule

• Use the atomic orbital diagrams to determine if an element is paramagnetic or diamagnetic

Be able to write **electron configurations for *ions* formed by elements #1-56**.

• Account for electrons gained or lost

• Write electron configurations using full notation and core notation

• Know Representative Elements usually form ions that are isoelectronic w/ a Noble Gas

• Show that each sublevel is slightly higher in energy than the previous sublevel

• Note that transition metals lose their s electrons before their d electrons.

Chapter 6: Ionic Bonds

ionic bond: electrostatic attraction holding cations and anions together in ionic compound

core electrons: innermost electrons in filled shells

valence electrons: outermost electrons

– element's group number = # of valence electrons

Periodic Trends:

Know definitions for the following: atomic radius, ionization energy, electron affinity

• **Atomic radius and Metallic Character:**

- Increases down a group
- Decreases left to right across a period

• **1st Ionization Energy & Electron Affinity:**

- Decreases down a group
- Increases left to right across a period

Know **trends for radii:** cation < neutral atom < anion

Acidic and Basic Oxides:

– Identify nonmetal oxides are generally acidic oxides while metal oxides are generally basic oxides

Ionic Compounds

– exist as 3D network of ions

– requires a large amount of energy to break all bonds to melt $\rightarrow$ have high melting points

– Relative strength of an ionic bond given by:

- Higher charge $\rightarrow$ stronger bond
- Shorter distance b/w ions $\rightarrow$ stronger bond

– Use Coulomb's Law to predict which ionic compound in a given pair will have the higher melting point

Chapter 7: Covalent Bonds, Lewis Structures, Molecular Geometry

covalent bond: sharing of electrons between two nonmetal atoms

• coordinate covalent: one atom donates both electrons to make the bond

• **polar covalent**: unequal sharing of electrons by 2 atoms with different electronegativities

• nonpolar covalent: equal sharing of electrons by two atoms with equal electronegativities

electronegativity: ability of an atom in a bond to draw electrons to itself

– Know F is most electronegative, further away from F, less electronegative; H between B and C.

octet rule: atoms bond such that each has 8 electrons, except H only needs 2 electrons.

Draw **Lewis (Electron Dot) Structures** for Molecules and Polyatomic Ions

– For ternary oxyacids, oxygens around the central atom, and one hydrogen bonded per oxygen atom

– Be able to draw **Lewis electron dot formulas** for atoms and ions to show # of valence electrons

– Account for electrons gained or lost for ions.

resonance structures: two or more structures representing a single molecule that cannot be described fully with only one Lewis structure

– Recognize molecules requiring resonance structures

Exceptions to the Octet Rule

- Incomplete octet
  - Be and B can have less than 8 e^-s
- N can have an odd-number of e^-s
- Elements in the third period and lower can have more than 8 electrons or an expanded octet for

bond length: distance b/w nuclei of 2 bonded atoms

- **shorter** the bond → the **stronger** the bond
- single bonds are the longest
- double bonds a the next shortest
- triple bonds are the shortest

Valence-Shell Electron Pair Repulsion (VSEPR)

- repulsions between electrons give rise to the shapes or molecular geometries of molecules
- A=central atom, B=outer atom(s), E=lone pair(s) on central

AB_2 → linear, 180° bond angle
AB_3 → trigonal planar, 120° bond angle
AB_4 → tetrahedral, 109.5° bond angle
AB_5 → trigonal bipyramidal, 90° & 120° angles
AB_6 → octahedral, 90° bond angle

Variation on trigonal planar shape
AB_2E → bent, <120° bond angle

Variations on tetrahedral shape
AB_3E → trigonal pyramidal, <109.5° bond angle
AB_2E_2 → bent, <109.5° bond angle

Valence Bond Theory

- Recognize that two atoms form a bond because they share electrons to become more stable

Polarity

- Use the shape & electronegativity differences to determine if a molecule is polar or nonpolar

Identify bonds as **sigma** and **pi bonds** using the Lewis electron dot formula of a molecule

- **sigma bond** (σ): bond formed from direct overlap of two orbitals
  - formed by single bonds and first bond for double and triple bonds
- **pi bond** (π): bond formed from sideways or indirect overlap of two orbitals
  - formed by the second and third bonds in double and triple bonds, respectively

Hybrid Atomic Orbitals

- Identify the hybrid atomic orbital for the central atom(s) in a molecule using the shape of the molecule

Shape	Atomic Hybrid Orbitals
linear (where central atom has an octet or incomplete octet)	sp
trigonal planar or **bent** (with <120° ∠)	sp^2
tetrahedral, **trigonal pyramidal**, or **bent** (with <109.5° ∠)	sp^3
trigonal bipyramidal, *see-saw, T-shape, or linear (with expanded octet)*	sp^3d
octahedral, *square pyramidal, or square planar*	sp^3d^2

Chapter 8: Thermochemistry

- **heat:** energy transferred from a hotter substance to a cooler substance
- **kinetic energy (KE):** $KE = \frac{1}{2}mv^2$
  - energy associated with an object's motion
  - Be able to solve kinetic energy problems!
- **potential energy:** energy associated with an object's position or its chemical bonds
- **joule (J):** $1\ J = \frac{1\ kg \cdot m^2}{s^2}$ and **1 kJ=1000 J**
- Distinguish between **system** and **surroundings**

- **heat of reaction ($q_{reaction}$):**
  - endothermic reaction: $q_{reaction}$ is +
  - exothermic reaction: $q_{reaction}$ is –
- Under **constant pressure** (like atmospheric pressure), $q_{reaction} = \Delta H$
  - for **endothermic** reactions, **ΔH is +**
  - for **exothermic** reactions, **ΔH is –**
- **Solve various problems involving the following:**
  - **heat capacity** (in **J/°C**): amount of heat to raise temp. of given amount by 1°C.
  - **molar heat capacity** (in **J/mol·°C**): heat capacity per mole of substance

q = (heat capacity) (ΔT)

  - **specific heat** (in **J/g·°C**): amount of heat to raise temp. of 1 gram of substance by 1°C.

q = (mass) (specific heat) (ΔT)

- **Thermochemical Equations** indicate the enthalpy changes as well as reactants and products
  - Be able to solve stoichiometry problems involving enthalpy changes

Guidelines for Thermochemical Equations

1. Magnitude of ΔH is proportional to the amount of reactants and products in the equation.
2. ΔH is the opposite sign for the reverse reaction.
3. If the coefficients in an equation are multiplied by a factor *n*, ΔH must able be multiplied by *n*.

Hess' Law: The value of ΔH for a reaction is the same whether it occurs in one step or several steps: $\Delta H = \Delta H_1 + \Delta H_2 + \Delta H_3 + \ldots$

- Be able to solve Hess' Law problems

Standard Enthalpy of Reaction (ΔH°_{rxn}):

- Enthalpy change associated with a reaction at standard state conditions

Standard Enthalpy of Formation (ΔH°_f):

- Enthalpy change for the **formation of 1 mole** of a **compound** from its **elements** in their **standard states.**

Solve problems using ΔH°_{rxn} and ΔH°_f:

- Be able to calculate ΔH°_{rxn} given ΔH°_f for all the reactants and products or
- Be able to calculate ΔH°_f for one reactant or product given ΔH°_{rxn} and ΔH°_f for the other reactants and products
-

bond energy: energy required to break a bond

- always positive
- **Use bond energies to calculate a reaction's ΔH**

Chapter 9: Gases

- Know the physical properties of gases
- Gas pressure and Atmospheric pressure
  - Standard atmospheric pressure: 760 torr (1 atm) at 0°C
  - Be able to convert between units of pressure: 1 atm ≡ 760 torr ≡ 760 mmHg = 101.325 kPa
  - Recognize how changes in temperature, pressure and number of moles of gas affect gas pressure
- Recognize volume increases with the formation of gases in a chemical reaction
- Solve for a variety of problems involving gases
  - Use ideal gas law (PV=nRT) to solve for P, V, n, or T. $R=\frac{0.0821\ L\cdot atm}{mol\cdot K}$ will be given.
  - Recognize that STP is standard temperature and pressure, defined as 0°C and 1 atm.
  - Use the ideal gas law and unit analysis to solve for density, molar mass, and other properties.
  - Solve stoichiometry problems using 22.4L/mol at STP and $n=\frac{PV}{RT}$ for other conditions.
  - Use volume-volume ratios when the reaction involves only gases at the same temperature and pressure
- **Dalton's Law of Partial Pressure:**
  - Use Dalton's Law ($P_{total} = P_1 + P_2 + P_3 + \ldots$) to solve for total pressure or the partial pressure of one or more gases in a mixture

- **Mole fraction** (X_A):

$$X_A = \frac{\#\text{ of moles of A}}{\text{total }\#\text{ of moles in mixture}}$$

- **Partial pressure** (P_A): $P_A = X_A P_{total}$
  Solve for mole fraction, total pressure, or partial pressure given other information

- **Kinetic Theory of an Ideal Gas**
  - Know the kinetic theory of gases regarding particle volume, particle motion, attraction, collisions, and the average kinetic energy of each particle being proportional to the temperature.
  - Molecular Speeds
    - Recognize that molecular speed is inversely related to the molar mass of a gas
      → The higher the molar mass of a gas, the slower it moves.

- Know about Diffusion and Effusion for gases

Chapter 10: Intermolecular Forces

Describe solids, liquids, and gases in terms of kinetic energy and molecular motion
- Know properties of liquids and solids

Intermolecular Forces: attraction between 2 different molecules in a liquid or solid
- Be able to identify intermolecular forces experienced by molecules in a given sample

Dispersion (Induced-Dipole or London) forces: attraction between molecules with instantaneous dipole resulting from electrons shifting
- strongest type of intermolecular force between **nonpolar** molecules
- becomes stronger with more electrons
  → the larger the nonpolar molecule → the stronger the dispersion forces

Dipole-dipole forces: attraction between polar molecules w/o H–O, H–N, or H–F bond
- Stronger than London forces because polar molecules have permanent dipoles

Hydrogen Bonds: attraction between polar molecules containing H–O, H–N, or H–F bond
- Strongest type of intermolecular force

Note that ionic and covalent bonds are stronger than any intermolecular force, even hydrogen bonds.

Intermolecular Forces and Physical Properties
- Recognize how intermolecular forces affect physical properties like boiling point, vapor pressure, molar heat of vaporization, viscosity, and surface tension.
- Identify which substance has higher boiling point based on intermolecular forces

Intermolecular Forces vs. Ionic/Covalent Bonds
- Distinguish between intermolecular force between different molecules and the bonds (ionic, polar covalent, or nonpolar covalent) holding atoms or ions together in a compound
- Identify type of bond broken for a chemical reaction or a change in physical state
  - e.g. hydrogen bonds are broken to boil water; ionic bonds are broken to melt NaCl; nonpolar covalent bonds are broken and polar covalent bonds are made when hydrogen gas and oxygen gas react to form water.

Structure and Properties of Water
- Recognize the density of ice is lower than the density of liquid water
- Because water molecules have a tetrahedral arrangement in ice, there are hexagonal holes in ice at the molecular level → why snowflakes have hexagonal (six-sided) symmetry

Crystalline Solids: w/ regular, geometric pattern
- Identify a solid as ionic, molecular (including network covalent solids), or metallic given its formula or its properties

Know terms for all phase transitions.
- evaporation (vaporization): liquid → vapor
- condensation: vapor → liquid
  - Cooling or applying pressure to gas → liquid
- sublimation: solid → vapor
- deposition: vapor → solid
- freezing: liquid → solid
- melting (or fusion): solid → liquid

normal melting point (m.p.) or boiling point (b.p.) always at ***1 atm***

heating curve: shows the physical state of a substance under certain temperatures and pressures

- Be able to use a heating curve to identify the phase changes a sample undergoes when it goes from one temperature to another.
- Recognize the features of a heating curve: melting and boiling points, the regions where the substance is solid only, solid+liquid, liquid only, liquid+gas, gas only.

phase diagrams: summarize the physical state of a substance at any temperature and pressure

- Identify the physical state of a substance at a given temp. or pressure using a phase diagram
- Know definitions of the following terms: triple point, critical temperature, critical pressure

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Chapter 11: Solutions

- Know definitions for the following terms: solution, solute, solvent, unsaturated, saturated, supersaturated, miscible/immiscible versus soluble/insoluble, colligative property, molality
- Know the effects of **temperature and pressure** on the **solubility of a solid or gas** in solution
- **"Like dissolves like" rule**
  - Polar solids and liquids will only dissolve in (or mix with) polar solvents
  - Nonpolar solids and liquids will only dissolve in (or mix with) nonpolar solvents
  - Polar solids and liquids NEVER dissolve or mix with nonpolar solvents
  - Nonpolar solids and liquids NEVER dissolve or mix with polar solvents
  - Use the Solubility Rules to determine which ionic compounds are soluble in polar solvents
  - Ionic compounds NEVER dissolve in nonpolar solvents
- Be able to use the **"like dissolves like" rule** to determine if a substance is soluble/insoluble/miscible/immiscible in a given solvent
- Recognize the following colligative properties: freezing point depression, boiling point elevation, vapor pressure lowering, osmotic pressure

Be able to combine various concepts to solve a variety of problems.

You will be given the usual Periodic Table with constants, Solubility Rules, Activity Series and Active Metals.

A scantron form will be provided. You will need to bring a #2 pencil and a basic scientific (non-programmable) calculator.