CHEM 1110 – Fall 2021
Midterm Exam 3 Study Guide (Ch. 6-8)

This study guide is meant to provide only the barest direction as you study. Try to find practice problems from the textbook (both in the chapter text and in the end-of-chapter questions) rather than just relying on this guide. Note that most tables and equations will not be provided here or on the exam. You can find them in your textbook.

Chapter 6 – Chemical Reactions: Mole and Mass Relationships

- Calculating molecular weights and formula weights
- Calculating moles from mass, and mass from moles
- Stoichiometric ratios between reactants and products
- Calculating theoretical yield
- Limiting reactant problems
  - Which is the limiting reactant?
  - How much excess reactant is left over?
  - What is the percent yield?

Chapter 7 – Chemical Reactions: Energy, Rates, and Equilibrium

- Energy is released by forming bonds and energy is required to break bonds
- Finding enthalpies of reaction from bond enthalpies
  $\Delta H_{rxn} \approx \sum Bonds\ Broken - \sum Bonds\ Formed$
- Find the enthalpy of reaction for the following reaction:
  $H_2(g) + Cl_2(g) \longrightarrow 2\,HCl$
- Answer: $\Delta H_{rxn} \approx -43\,\frac{kcal}{mol}$
- Question: If $5.00\ g$ of H_2 react with excess Cl_2 to produce HCl, how much heat is released?
- Answer: $107\ kcal$
- Exothermic vs Endothermic reactions
- Free energy and spontanaety (negative $\Delta G_{rxn} is spontaneous$)
  $\Delta G_{rxn} = \Delta H_{rxn} - T\Delta S_{rxn}$

- Suppose a reaction has $\Delta H_{rxn} > 0$ and $\Delta S_{rxn} > 0$. Will this reaction be spontaneous at high, low, all, or no temperatures?
- Answer: High temperature
- Rates of chemical reactions are controlled by three factors:
  - The frequency of collisions - Concentration of reactants
  - The energy of collisions - Must be greater than activation energy. Temperature dictates average energy
  - Reaction pathway - A catalyst could speed a reaction
- True/False Increasing the temperature will decrease the reaction rate for exothermic reactions.
- Answer: False. Increasing the temperature always increases the reaction rate.
- Reaction coordinate diagrams
- Some reactions are *reversible*. They reach an equilibrium rather than going all the way to products
- The equilibrium conditions are governed by the equilibrium constant: $K_{eq} = \frac{[Products]^m}{[Rreactants]^n}$
- Give the equilibrium constant for the following reaction:

  $PO_4^{3-}(aq) + 2\,H_2O(l) \longleftrightarrow H_2PO_4^-(aq) + 2\,OH^-(aq)$
- Answer: $K_{eq} = \frac{[H_2PO_4^-][OH]^2}{[PO4^{3-}]}$ (note that the water is absent because it is (l))
- Le Châtelier's Principle:
  - Shifting in response to adding or removing reactants or products
  - Shifting in response to changes in pressure
  - Shifting in response to changes in temperature

Chapter 8 – Gases, Liquids, and Solids

- ΔH, ΔS, and ΔG for all phase changes
- Intermolecular forces
  - London Dispersion Forces
  - Dipole-dipole Forces
  - Hydrogen Bonds
- Question: Is methanol (CH_3OH) capable of hydrogen bonding?
- Answer: Yes, it has both a H-bond donor and acceptor, so it can hydrogen bond

- Know the assumptions that make up the kinetic molecular theory of gases
- Pressure an the many different units for expressing it
- Gas laws: $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$
- Question: A balloon has $V = 3.21\ L$ at $P = 0.82\ atm$ and $T = 298\ K$. What temperature would give the balloon a volume of $V = 2.95\ L$ at $P = 0.67\ atm$?
- Answer: $T = 224\ K$
- Ideal Gas Law: $PV = nRT$
- Dalton's law of partial pressures
- Vapor pressures, boiling points, and Henry's law
- Viscosity and surface tension
- Different types of crystalline solids and amorphous solids
- Heats for phase changes: $q = m\Delta H$
- Question: How much heat is absorbed when $12.3\ g$ of dry ice sublimate in your home-made root beer? ($\Delta H_{sub} = 571\frac{J}{g}$)
- Answer: $7.02\ kJ$