

# CHEM 1220 Lecture Notes

## OpenStax Chemistry 2e

Matthew Rowley

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## COURSE ADMINISTRATIVE DETAILS

- My office hours
- Intro to my research
- Introductory Quiz
- Grading details
  - Exams - 40, Final - 15, Online Homework - 15, Book Homework - 15, Quizzes - 15
  - Online homework
  - Frequent quizzes
- Importance of reading and learning on your own
- Learning resources
  - My Office Hours
  - Tutoring services - <https://www.suu.edu/academicsuccess/tutoring/>
- Show how to access Canvas
  - Calendar, Grades, Modules, etc.
  - Quizzes
  - Textbook
- Introduction to chemistry
  - Ruby fluorescence
  - Levomethamphetamine
  - Rubber band elasticity
  - Structure of the periodic table
  - Salt on ice and purifying hydrogen peroxide

## CHAPTER 0

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1210 REVIEW

There is a whole semester of material from 1210, and these are only the topics which are *most* important for success in 1220

- Composition of atoms and ions (protons, neutrons and electrons)
- Chemical formulas and names
  - Formulas and molar masses
  - Polyatomic ion names
  - Naming ionic compounds
  - Naming binary molecular compounds
  - Naming acids
- Balancing molecular equations
- Solubility rules
- Fundamentals of acid/base chemistry
- Measurements in chemistry
  - Converting from measurements to moles and back
  - Stoichiometry and predicting amounts
  - Limiting reactants
- Enthalpy of reaction and heat equations
- Lewis structures

**CHEM 1210 Review Quiz**

## CHAPTER 10

## LIQUIDS AND SOLIDS

## 10.1 Intermolecular Forces

- Many physical properties of solids, liquids, and gases can be explained by the strength of attractive forces between particles (Figure 10.5)
- Phase changes happen due to the interplay between kinetic energy and intermolecular forces (Figure 10.2)
- Pressure can also play a role in phase changes, as discussed later
- These *intermolecular forces* come in different varieties
  - Dispersion Forces Non-polar molecules, impacted by polarizability, molecular weight, and surface area
    - \* Dominant in non-polar molecules
    - \* Created by induced dipoles (Figure 10.6)
    - \* Impacted by polarizability (Table 10.1)
    - \* Impacted by molecular weight (hydrocarbons from methane to wax)
    - \* Impacted by molecule shape (Figure 10.7 compares the boiling points of pentane isomers)
  - Dipole-Dipole Forces
    - \* Dominant in polar molecules
    - \* Results from attraction between permanent dipoles (Figure 10.9)
  - Hydrogen Bonding
    - \* Dominant only in molecules capable of hydrogen bonding
    - \* Must contain a hydrogen-donor atom (H attached to N, O, or F)
    - \* Must contain a hydrogen-acceptor atom (lone pair of electrons attached to N, O, or F)
    - \* Hydrogen bonds are more than just particularly strong dipole-dipole forces. They have strong directionality according to VSEPR
    - \* Figures 10.10, 10.14, and other figures on the Internet show water, DNA, and proteins all organized by hydrogen bonds
    - \* Figures 10.11 and 10.12 illustrate how much hydrogen bonds exceed dipole-dipole forces in strength

## 10.2 Properties of Liquids

- Viscosity is a fluid's resistance to flow

- We intuitively know that both water and honey flow...but at very different rates
- Viscosity is proportional to the strength of intermolecular forces (high IF = high viscosity)
- As temperature increases, kinetic energy is able to overcome intermolecular forces and viscosity decreases
- Table 10.2 gives the viscosities of some common substances (note the unusual units!)
- Surface tension is a force which minimizes a fluid's surface area
  - Cohesive vs. adhesive forces
  - Bulk molecules have lower energy than surface molecules due to being *surrounded* by cohesive forces (Figure 10.16)
  - Figure 10.17 illustrates a waterbug supporting itself on water surface tension
  - Surface tension is often in conflict with gravity and other forces, making most liquids rounded but not perfect spheres
  - Surface tension is proportional to intermolecular forces (Table 10.3)
  - Surface tension can be strongly affected by addition of certain solutes, called surfactants
- Capillary action is a force between a fluid and narrow channels or capillaries of solid materials
  - Due to adhesive forces with the solid, liquids will be drawn up (or, less often, pushed down) a capillary
  - Figure 10.19 shows how paper towels are made to maximize capillary action, so they soak up water-based spills
  - The top of the liquid (called the meniscus) will curve differently depending on the relative strength of cohesive and adhesive forces (Figure 10.18)
  - Figure 10.20 shows capillary action in a variety of situations, including capillary repulsion
  - Remember that when measuring volumes, convention is to read the *bottom* of the meniscus regardless of how it curves
  - Don't worry about the formula given here

### Quiz 10.1 - Intermolecular Forces and Liquid Properties

#### Homework 10.1

- 10.11: Predicting trends in boiling points
- 10.21: Identifying intermolecular forces
- 10.25: Affect of temperature on viscosity

### 10.3 Phase Transitions

- Vaporization and condensation are the transitions between liquid and gas phases
  - The enthalpy of vaporization ( $\Delta H_{vap}$ ) is the energy required to transition from liquid to gas phase
  - Enthalpy of condensation is the opposite  $\Delta H_{con} = -\Delta H_{vap}$
  - In a closed volume, these processes will reach a *dynamic equilibrium*
  - The partial pressure of the liquid at this equilibrium state is called its *vapor pressure* (Figure 10.22)
  - Higher intermolecular forces lead to lower vapor pressures
  - Higher temperatures increase the vapor pressure due to increased kinetic energy (Figure 10.23)
- Boiling points
  - Figure 10.24 shows vapor pressure curves and the normal boiling points of several liquids
  - Boiling points generally depend on the pressure (pressure cookers, boiling water to freezing, etc.)
  - The Clausius-Clapeyron equation defines these curves (Note the rearrangements I've made)
 
$$P = Ae^{-\Delta H_{vap}/RT} \quad \ln P = -\frac{\Delta H_{vap}}{RT} + \ln A \quad \ln \left( \frac{P_2}{P_1} \right) = -\frac{\Delta H_{vap}}{R} \left( \frac{1}{T_2} - \frac{1}{T_1} \right)$$
- Fusion (melting), freezing, sublimation, and deposition all have their enthalpies and transition temperatures
- These enthalpies are state functions, such that  $\Delta H_{sub} = \Delta H_{fus} + \Delta H_{vap}$  (Figure 10.28)
- Heating and Cooling curves
  - When heat is added to a system, it will either cause a phase change, or a change in temperature
  - For phase changes,  $q = n\Delta H_{change}$
  - For temperature changes,  $q = mc\Delta T$ , where  $c$  is the specific heat for that substance and phase
  - Sometimes  $\Delta H_{change}$  is given as a -per gram value, and sometimes  $c$  is given as a -per mole value, but usually not :(
  - Figure 10.29 shows a typical heating curve (Work example 10.10 in the text)

#### Quiz 10.2 - Heating Curves

#### Homework 10.2

- 10.31: Temperature during a phase transition
- 10.39: Definition of normal boiling point
- 10.51: Heating curve problem

## 10.4 Phase Diagrams

- The stable phase at different temperatures and pressures is best illustrated with a phase diagram (Figures 10.30, 10.31)
- We can tell at a glance what transitions might occur as we increase or decrease either the temperature or pressure
- Note that at some pressures, sublimation may occur instead of fusion
- The triple point is a unique point where liquid, solid, and gas can all exist at equilibrium (contrast with a glass of icy water on a humid day)
- The critical point is where the distinction between liquid and solid phases disappears
- Figure 10.34 shows the phase diagram of  $\text{CO}_2$
- Supercritical fluids exhibit some interesting properties, and are often great solvents (Nile Blue Youtube video)
- Critical points vary widely depending on the intermolecular forces, and other factors (Table in text)

### Quiz 10.3 - Phase Diagrams

### Homework 10.3

- 10.55: Trajectories on a phase diagram
- 10.57: Determining state on a phase diagram
- 10.63: Identifying phases on a blank phase diagram

## 10.5 The Solid State of Matter

- Solids can be divided into *crystalline* and *amorphous* based on their structure at atomic scales
- Figure 10.37 shows the difference generally, Figure 10.38 shows crystalline and amorphous  $\text{SiO}_2$
- Amorphous solids will not exhibit a sharp fusion transition temperature, but will instead grow soft and malleable over a temperature range
- Crystalline solids are diverse but always show long-range repeating order in their structure
  - Ionic solids (Figure 10.39) have high melting points, cleave along planes, and conduct electricity only in the liquid phase
  - Metallic solids (Figure 10.40) have mostly high melting points, are malleable and ductile, and conduct electricity and heat well

- Covalent network solids (Figure 10.41) have very high melting points and are electrical insulators
- Molecular solids (Figure 10.42) Have low to very low melting points and are electrical insulators
- Crystalline solid properties are summarized in Table 10.4
- Even crystalline solids do not have perfect structure. Various types of defects are illustrated in Figure 10.45

### Quiz 10.4 - Classifying Solids

### Homework 10.4

- 10.69: Classify solids by formulas
- 10.71: Classify solids by properties

## 10.6 Lattice Structures in Crystalline Solids

- The structure of a crystalline solid is represented by a *unit cell*, the smallest repeatable unit of the structure
- Sometimes this microscopic structure is evidently manifested on macroscopic scales, but sometimes it isn't
- Unit cells are defined by lattice points that often lie at the center of certain atoms, and the cell edges often cut atoms in half, quarter, etc.
- Unit cells of metals
  - For metals, we should keep track of the quantity of atoms in a unit cell, the coordination number, and the relationship between the atomic radius and unit cell edge length
  - Simple cubic (Figure 10.49) 1 atom, Coordination=6,  $l = 2r$
  - Body-centered cubic (Figure 10.51) 2 atoms, Coordination=8,  $l = \frac{4}{\sqrt{3}}r$
  - Face-centered cubic (Figure 10.52) 4 atoms, Coordination=12,  $l = \sqrt{8}r$
  - Figure 10.54 shows hexagonal closest packed and cubic closest packed structures
  - Find the radius of a gold atom, which has fcc structure and a density of  $19.283\text{g}/\text{cm}^3$  ( $136\text{pm}$ )
  - Find the density of polonium, which has sc structure and an atomic radius of  $140\text{pm}$  ( $9.20\text{g}/\text{cm}^3$ )
  - Figure 10.56 shows many non-cubic structures which are common as well
- Unit cells of ionic compounds
  - Anions are generally larger than cations, so ionic lattice points are generally the centers of anions
  - Cations occupy holes in the anionic lattice (Figures 10.57 and 10.58)



- Unit cells of ionic structures share names with the metallic cells but look different because of the cations
- Simple cubic (Figure 10.59)
- Face-centered cubic (rock salt structure) (Figure 10.60)
- Zinc blende (Figure 10.61)
- Find the ionic bond length for NaCl which has rock salt structure and density of  $2.17 \text{ g/cm}^3$  ( $l = 564 \text{ pm}$ )
- Crystal structure is determined through X-ray crystallography
  - X-rays reflected off a crystal surface can combine destructively or constructively to produce an interference pattern (Figure 10.63)
  - The X-rays will take different pathlengths depending on the angle of the X-ray beam and the crystal lattice constant (Figure 1.64)
  - An experimental setup and actual diffractogram are shown in Figures 10.65 and 10.66
  - We have a powerful X-ray instrument here at SUU

### Quiz 10.5 - Unit Cells

### Homework 10.5

- 10.77: Coordination number
- 10.81: Density from lattice constant
- 10.85: Packing efficiency and density

CHAPTER 11

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SOLUTIONS AND COLLOIDS

**11.1 The Dissolution Process**

**11.2 Electrolytes**

**11.3 Solubility**

**11.4 Colligative Properties**

**11.5 Colloids**

CHAPTER 12

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KINETICS

**12.1 Chemical Reaction Rates**

**12.2 Factors Affecting Reaction Rates**

**12.3 Rate Laws**

**12.4 Integrated Rate Laws**

**12.5 Collision Theory**

**12.6 Reaction Mechanisms**

**12.7 Catalysis**

CHAPTER 13

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FUNDAMENTAL EQUILIBRIUM CONCEPTS

- 13.1 Chemical Equilibria**
- 13.2 Equilibrium Constants**
- 13.3 Shifting Equilibria: Le Châtelier's Principle**
- 13.4 Equilibrium Calculations**

CHAPTER 14

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ACID-BASE EQUILIBRIA

- 14.1 Brønsted-Lowry Acids and Bases
- 14.2 pH and pOH
- 14.3 Relative Strengths of Acids and Bases
- 14.4 Hydrolysis of Salts
- 14.5 Polyprotic Acids
- 14.6 Buffers
- 14.7 Acid-Base Titrations

CHAPTER 15

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EQUILIBRIA OF OTHER REACTION CLASSES

**15.1 Precipitation and Dissolution**

**15.2 Lewis Acids and Bases**

**15.3 Coupled Equilibria**

CHAPTER 16

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THERMODYNAMICS

**16.1 Spontaneity**

**16.2 Entropy**

**16.3 The Second and Third Laws of Thermodynamics**

**16.4 Free Energy**

CHAPTER 17

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ELECTROCHEMISTRY

- 17.1 Review of Redox Chemistry**
- 17.2 Galvanic Cells**
- 17.3 Electrode and Cell Potentials**
- 17.4 Potential, Free Energy, and Equilibrium**
- 17.5 Batteries and Fuel Cells**
- 17.6 Corrosion**
- 17.7 Electrolysis**



CHAPTER 21

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NUCLEAR CHEMISTRY

- 21.1 Nuclear Structure and Stability**
- 21.2 Nuclear Equations**
- 21.3 Radioactive Decay**
- 21.4 Transmutation and Nuclear Energy**
- 21.5 Uses of Radioisotopes**
- 21.6 Biological Effects of Radiation**

## ERRATA