As you work through each chapter, you should be able to:

**Chapter 10 – Solids and Liquids**
1. Differentiate between the various types of intermolecular forces that hold solid and liquid molecules close together.
2. Identify the types of particles occupying the lattice points of the crystal in a crystalline structure.
3. Explain thermal and electrical conductivity of metals using the “sea of electrons” model.
4. Differentiate between the structure and properties of network, metallic, group 8A, molecular, and ionic solids.
5. Calculate the heat changes associated with changes in state.
6. Understand the relationship between vapor pressure and phase change.
7. Explain the changes that are occurring at each point along a heating or cooling curve.
8. Draw and understand phase diagrams for one-component systems.
9. Know the significance of the critical point and triple point.
10. Understand the unique nature of the solid-liquid line on the water phase diagram.

**Chapter 11 – Properties of Solutions**
1. Describe composition of the different types of solutions.
2. Calculate composition of a solution using molarity, molality, mass percent, volume percent, and mole fractions.
3. Calculate the heat of solution formation and enthalpy of hydration.
4. Explain how structure, pressure, and temperature influence solution solubility in ideal and non-ideal solutions.
5. Use Raoult’s law to determine the vapor pressure of solutions.

**Lab Experience:**
1. Evaporation and Intermolecular Attractive Forces
2. Determination of the Heat of Solution

**Assignments:**
- HW1 Ch 10 Handout
- Hw2 Ch 10 Text Problems
- HW3 Ch 10 Quiz and Reading Notes Check
- HW4 Ch 11 Handout
- HW5 Ch 11 Text Problems
- HW6 Ch 11 Quiz and Reading Notes Check

**Week of January 5th**

<table>
<thead>
<tr>
<th>Day</th>
<th>Concepts</th>
<th>Class Activities</th>
<th>Assignments</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mon</td>
<td>Ch 10</td>
<td>-Ch 10 Notes</td>
<td>Assign:</td>
</tr>
<tr>
<td>1 – 6</td>
<td></td>
<td></td>
<td>-HW 1, HW 2, HW 3</td>
</tr>
<tr>
<td>Blk 1</td>
<td>Ch 10</td>
<td>-Ch 10 Notes</td>
<td>Assign:</td>
</tr>
<tr>
<td></td>
<td>Lab</td>
<td>-Evaporation and Intermolecular Forces</td>
<td>-Evaporation and Intermolecular Forces Pre-lab &amp; Lab</td>
</tr>
<tr>
<td></td>
<td></td>
<td>-Demos</td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td>-Pre-lab</td>
<td></td>
</tr>
<tr>
<td>Blk 2</td>
<td>Ch 10</td>
<td>Evaporation and Intermolecular Forces Lab</td>
<td></td>
</tr>
<tr>
<td></td>
<td>Lab</td>
<td>-Collect Data, begin reports</td>
<td></td>
</tr>
<tr>
<td>Fri</td>
<td>Ch 10</td>
<td>-Finish Ch 10 Notes</td>
<td>Due</td>
</tr>
<tr>
<td></td>
<td></td>
<td>-Demos</td>
<td>• HW1</td>
</tr>
<tr>
<td></td>
<td></td>
<td>-Work time on Lab reports</td>
<td></td>
</tr>
</tbody>
</table>
### Chapter 10 – Liquids and Solids Notes

**States of Matter**
- Substances can exist as solids, liquids, and gases at room temperature.
- Each state has different physical properties.
- Look at H$_2$O:
  \[ \Delta H_{\text{fus}} = 6.02 \text{ kJ/mol} \]
  \[ \Delta H_{\text{vap}} = 40.7 \text{ kJ/mol} \]

**Liquids and Solids**
- Liquids and solids show many similar properties.
- The properties of solids and liquids are considerably different than those of gases.
- Properties of solids and liquids can be attributed to intermolecular forces.

**Intermolecular Forces**
- Intermolecular forces occur between substances (as opposed to intramolecular forces that exist between atoms in a compound).
- Many types of intermolecular forces exist between substances.

**Dipole-Dipole Forces**
- Polar bonds behave as if they have a center of positive and negative charge
- Molecules with dipole moments can attract themselves electrostatically
- This electrostatic attraction between oppositely charged poles of a molecule is a dipole-dipole interaction
- These attractions are typically ~1% the strength of a covalent and ionic bond

**Hydrogen Bonding**
- Dipole-dipole interactions are stronger in bonds with hydrogen atoms.
- Hydrogen is very small and allows the molecules to come in very close proximity to each other.
- Hydrogen bonding accounts for the unusually high boiling points of covalent hydrides of the first elements in Groups 5A, 6A and 7A.

**London Dispersion Forces**
- All molecules, whether or not they are polar, exert forces on each other.
- London dispersion forces, often called induced dipole forces, account for nonpolar substances existing in the solid or liquid phase.
- Electrons in an electron cloud can induce a temporary dipole in a neighboring atom.
- This interaction is weak and short-lived but can be quite significant in larger atoms

### Week of January 12th

<table>
<thead>
<tr>
<th>Day</th>
<th>Concepts</th>
<th>Class Activities</th>
<th>Homework</th>
</tr>
</thead>
<tbody>
<tr>
<td>M</td>
<td>Ch 10 &amp; 11</td>
<td>-Chapter 11 Notes</td>
<td>Due</td>
</tr>
<tr>
<td>1 – 6</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>1st Block</td>
<td>CH 10 &amp; 11</td>
<td>-Ch 11 Notes -Chapter 10 Quiz</td>
<td>Due</td>
</tr>
<tr>
<td>R</td>
<td>Ch 11</td>
<td>-Ch 11 Notes -Demos</td>
<td>Due</td>
</tr>
<tr>
<td>1 – 6</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>F</td>
<td>Ch 10 &amp; 11</td>
<td>-Test Review time Ch 10 &amp; 11</td>
<td>Due</td>
</tr>
<tr>
<td>1 – 6</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

### Week of January 19th

<table>
<thead>
<tr>
<th>Day</th>
<th>Concepts</th>
<th>Class Activities</th>
<th>Homework</th>
</tr>
</thead>
<tbody>
<tr>
<td>M</td>
<td>No School:</td>
<td>Non-Contract Day</td>
<td></td>
</tr>
<tr>
<td>T</td>
<td>Ch 10 &amp; 11</td>
<td>- Test Review time Ch 10 &amp; 11 -Ch 11 Quiz</td>
<td>Due</td>
</tr>
<tr>
<td>1 – 6</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>F</td>
<td>Ch 10 &amp; 11</td>
<td>Finals (Chapter 10 &amp; 11 Exam) No retakes available since it is end of semester.</td>
<td>Due</td>
</tr>
<tr>
<td>5&amp;6</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
The Liquid State
- Liquids have:
  - Low compressibility
  - Lack of rigidity
  - High density compared with gases
- Forces of attraction in a liquid depend on where the molecule is in the liquid.
- Molecules on the interior interact with themselves, while molecules on the surface react with the liquids surroundings.

Adhesion vs. Cohesion
- Cohesion – forces among liquids
- Adhesion – forces between molecules and container
- Meniscus:
  - Concave – adhesive forces are stronger than cohesive forces
  - Convex – adhesive forces are weaker than cohesive forces

Surface Tension
- Surface tension – resistance of a liquid to increase surface area
- Polar molecules tend to form a concave meniscus when placed in a glass tube.
- Nonpolar liquids like mercury form a convex meniscus in a glass tube.

Capillary Action
- Capillary action – spontaneous rising of a liquid in a tube
- Exhibited by polar molecules that have an attraction to the capillary tube.

Viscosity
- Viscosity – measurement of a liquid’s resistance to flow
- Larger molecules exhibit higher viscosity

Introduction to Structure and Types of Solids
- Two main types of solid structures
- Crystalline solids – highly ordered arrangement of components (ionic salts)
- Amorphous solids – considerable disorder in their structure (glass)

Crystalline Solids
- Lattice:
  - 3-D system of points (atoms, molecules, and ions)
- Unit Cell:
  - Smallest repeating unit of a lattice
  - Three common unit cells include – simple cubic, body-centered cubic, and face-centered cubic
- Crystalline structures are determined using X-ray diffraction.

Types of Atomic Solids
- Ionic solids – ions at lattice points
- Molecular solids – discrete covalently bonded molecules at lattice points
- Atomic solids – lattice points are elements
  - Metallic – delocalized non-directional covalent bonding
  - Network – bond is strong, localized, and directional
  - Group 8A – Noble gases held together by London Forces

Metallic Crystal Structures
- Metallic crystal structures follow the closest packing model.
- The closest packing model assumes metal atoms are uniform, hard spheres.
- These spheres then pack together in a manner that most efficiently uses the available space.

Bonding Model for Metals
- Metals are characterized by having:
  - High thermal conductivity
  - High electrical conductivity
  - Malleable
  - Ductile
  - High melting point
- These properties are a result of the non-directional covalent bonding found in metallic crystals.
Bonding Continued
  • Metal ions exist spread out in a closest packed structure with a “sea of electrons” in between them.
  • This accounts for a metals ability to be bent and stretched without breaking.
  • It is difficult to separate metal atoms, but easy to move them.
  • Mobile electrons can also conduct heat and electricity

Network Atomic Solids
  • Large molecules contain strong directional bonds.
  • Do not efficiently conduct heat or electricity.

Carbon
  1. Diamond – hard colorless, nonconductor, insulator, tetrahedral arrangement with 4 other carbons
  2. Graphite – black, slippery conductor
    • Layers of six carbon rings (sp² hybridized)
    • Delocalized electrons in pi bonds
    • Strong bonding within the layers, but weak bonding between them

Silicon
  • Silicon is to geology as carbon is to biology. Different from carbon compound bonding although they are in the same group.
    – Carbon makes long C-C chains while silicon makes Si-O chains
  • Silica - SiO₂ is the fundamental compound
  • CO₂ uses sigma and pi bonding, but silicon is too large for that. Silicon satisfies its octet by forming single bonds with four oxygen atoms (quartz).
  • When cooled quickly, silica compounds can form amorphous glass.

Molecular Solids
  • Network solids – atoms in lattice positions act as one giant molecule
  • Some have discrete molecules at lattice points (H₂O, CO₂, S₈, P₄). Parts of structure have intramolecular attraction, but parts are held together through intermolecular attraction.

Ionic Solids
  • Stable, high-melting substances held together by strong electrostatic forces between oppositely charged ions.

Types and Properties of Solids

<table>
<thead>
<tr>
<th>Type of Solid:</th>
<th>Network</th>
<th>Metallic</th>
<th>Group 8A</th>
<th>Molecular</th>
<th>Ionic</th>
</tr>
</thead>
<tbody>
<tr>
<td>Structural Unit:</td>
<td>Atom</td>
<td>Atom</td>
<td>Atom</td>
<td>Molecule</td>
<td>Ion</td>
</tr>
<tr>
<td>Type of Bonding</td>
<td>Directional covalent bonds</td>
<td>Nondirectional covalent bonds</td>
<td>London dispersion forces</td>
<td>Polar: Dipole-dipole Nonpolar: London dispersion</td>
<td>Ionic</td>
</tr>
<tr>
<td>Typical Properties:</td>
<td>Hard, high melting point, insulator</td>
<td>Wide range of hardness and melting points, Conductor</td>
<td>Very low melting</td>
<td>Soft, low melting points, insulator</td>
<td>Hard, high melting point, insulator</td>
</tr>
<tr>
<td>Examples:</td>
<td>Diamond</td>
<td>Silver, iron, brass</td>
<td>Argon (s)</td>
<td>Water (s), Dry Ice (s)</td>
<td>Sodium chloride</td>
</tr>
</tbody>
</table>

Changes in State
  • The state of matter a substance exists in depends on its environment.
  • Three states of matter:
    – Solid
    – Liquid
    – Gas

Vapor
  • Vaporization (evaporation) is the endothermic process in which a liquid becomes a gas.
  • Heat of Vaporization:
    – Energy needed to vaporize 1 mole of liquid
    – Water has a high ΔH_vap = 40.7 kJ/mol, meaning it can absorb a lot of heat
Vapor Pressure

- Vapor – gas phase of a substance that exists as a liquid at 25°C and 1 atm.
- There is a dynamic equilibrium between a liquid and its vapor.
- A liquid/vapor system has reached equilibrium when the rate at which liquid molecules become vapor and vapor molecules become liquid is the same.
- Vapor pressure – pressure exerted by vapor at equilibrium

Measuring Vapor Pressure

- Vapor pressure can be measured using a simple barometer.
- \( P_{\text{atmosphere}} = P_{\text{vapor}} + P_{\text{Hg column}} \)
- \( P_{\text{vapor}} = P_{\text{atmosphere}} - P_{\text{Hg column}} \)
- Vapor pressure of a substance increases with temperature.

Graphical Analysis of Vapor Pressure vs. Temperature

- A plot of vapor pressure vs. temperature is nonlinear.
- \( \ln(P_{\text{vap}}) \text{ vs. } 1/T(\text{K}) \) is linear and follows \( y = mx + b \)
- \( y = \ln P_{\text{vap}}, x = 1/T, m = -\Delta H_{\text{vap}}/R, C = \text{a constant characteristic of the liquid}, R = 8.314 \text{ J/(Kmol)} \)
- Clausius-Clapeyron Equation:
  \[
  \ln P_{\text{vap}} = -\frac{\Delta H_{\text{vap}}}{R}(1/T) + C \quad \text{or} \quad \ln\left(\frac{P_{\text{vap,T1}}}{P_{\text{vap,T2}}}\right) = \frac{\Delta H_{\text{vap}}}{R}(1/T_2 - 1/T_1)
  \]
- This relationship is useful for determining either enthalpies of vaporization from vapor pressure measurements or vapor pressures at different temperatures.

Example

- The boiling point of acetone at 760. torr is 56.5°C. At an elevation of 5,300 ft the atmospheric pressure is 630. torr. What would be the boiling point of acetone (\( \Delta H = 30.2 \text{ kJ/mol} \)) at this elevation? What would be the vapor pressure of acetone at 25°C at this elevation?

Sublimation

- A solid to gas phase change is called sublimation (opposite is called deposition).
- This occurs because solids have vapor pressures.

Changes of State

- Changes in state can be represented using a heating curve (a plot of temperature vs. time).
- Solid \( \leftrightarrow \) Liquid = Heat of Fusion (\( \Delta H_{\text{fus}} \))
  - Point at which solid and liquid have identical vapor pressures
- Liquid \( \leftrightarrow \) Gas = Heat of Vaporization (\( \Delta H_{\text{vap}} \))
  - Point at which liquid and vapor have the same vapor pressure under conditions in which the system is under a constant pressure of 1 atm.
  - \( \Delta H_{\text{vap}} > \Delta H_{\text{fus}} \)

Heating Curves

- When heating or cooling a substance at a constant rate \( q = mc\Delta T \).
- At phase boundaries, the temperature remains constant until the entire substance has changed state. Only then can the substance be heated or cooled.
- Heating Curve for H\(_2\)O
Example

- A beaker contains 50.0 g of water and 25.0 g of ice, both at 0.0°C. How many kilojoules of energy is need to raise the final temperature of the water to 22.0°C? The specific heat capacity of liquid water is 4.18J/g°C and the enthalpy of fusion for ice is 6.02 kJ/mol.

Phase Diagrams

- Phase diagrams represent the phases of a substance as a function of temperature and pressure.
- Critical Temperature – temperature above which no applied pressure can liquefy a vapor
- Critical Pressure – pressure required to produce liquefaction at the critical temperature
- Critical Point – point at which the critical temperature and pressure are achieved
- Triple Point – point at which the vapor pressure of the solid, liquid, and gaseous phases are in equal

Normal Melting and Boiling Points

- Normal Melting Point – temperature at which the solid and liquid states have the same vapor pressure under conditions where the total pressure is 1 atmosphere
- Normal Boiling Point – temperature at which the vapor pressure of a liquid is exactly one atmosphere

Chapter 11 – Properties of Solutions

Solutions

- Most substances are mixtures.
- Solutions can be composed of solids, liquids, and gases.
- The composition of a mixture can be expressed in several ways:
  - Molarity
  - Mass percent
  - Mole fraction
  - Molality

Solution Composition

- Molarity (M):
  (moles of solute/liter of solution)
- Mass Percent:
  (mass of solute/mass of solution)x100%
- Mole Fraction (X):
  (moles of component)/(total number of moles)
- Molality (m):
  (moles of solute)/(kilogram of solvent)

Practice Problem

- A solution is prepared by mixing 1.00 g of ethanol (C₂H₅OH) with 100.0 g water to give a final volume of 101.0 ml. Calculate the molarity, mass percent, mole fraction, and molality of the ethanol in this solution.
The Energies of Solution Formation

- Solubility – ability to dissolve
- Three energy considerations:
  1. \( \Delta H_1 \) - Separating solute into individual components (expanding the solute)
  2. \( \Delta H_2 \) - Overcoming intermolecular forces in solvent to make room for solute (expanding the solvent)
  3. \( \Delta H_3 \) - Solute and solvent interact to form a solution

- \( \Delta H_1 \) and \( \Delta H_2 \) are endothermic
- \( \Delta H_3 \) is exothermic
- \( \Delta H_{\text{solution}} = \Delta H_1 + \Delta H_2 + \Delta H_3 \)

Like Dissolves Like

- The energies of solution formation can be used to understand the “like dissolves like” concept.
- Consider oil and water:
  1. It takes a lot of energy to overcome London Dispersion forces in a large oil molecule, therefore \( \Delta H_1 \) is large.
  2. It takes a lot of energy to overcome hydrogen bonding in water to make room for oil molecules, therefore \( \Delta H_2 \) is large.
  3. Oil is nonpolar and water is polar, therefore there is no real solvent solute interaction and \( \Delta H_3 \) is small.
  4. \( \Delta H_{\text{solution}} = \Delta H_1 + \Delta H_2 + \Delta H_3 \) = endothermic
  5. A large endothermic \( \Delta H_{\text{solution}} \) usually indicates immiscibility.
  6. Note: When \( \Delta H_{\text{solution}} \) is (+), but small, disorder tends to make the solute soluble (like NaCl \( \Delta H_{\text{solution}} = 3 \) kJ/mol).

Factors Affecting Solubility

- Structural Effect:
  1. Polar or nonpolar
- Pressure Effect:
  1. Negligible influence on liquids and solids
  2. Large influence on the solubility of gases
  3. Henry’s Law: \( P = kC \), \( P \) = partial pressure of solute, \( C \) = concentration, and \( k \) = constant characteristic of a particular solution
- Temperature Effect:
  1. Solubility usually increases with temperature
  2. The dependence of the solubility of a particular solute on temperature can only be determined experimentally

Henry’s Law Problem

- A certain soft drink is bottled so that at 25°C contains CO\(_2\) gas at a pressure of 5.0 atm over the liquid. Assuming that the partial pressure of CO\(_2\) in the atmosphere is 4.0 x 10\(^{-4}\) atm, calculate the equilibrium concentrations of CO\(_2\) in the soda both before and after the bottle is opened. The Henry’s law constant for CO\(_2\) in aqueous solution is 32 Lxatm/mol at 25°C.

The Vapor Pressure of Solutions

- A nonvolatile solute has no tendency to escape from solution into the vapor phase.
- The presence of a nonvolatile solute reduces the tendency of a solvent molecule to escape (solute decreases # of solvent molecules at the surface so fewer can escape).

Raoult’s Law

- For solutions containing nonvolatile solutes:
  \[ P_{\text{solution}} = X_{\text{solvent}} P^0_{\text{solvent}} \]
  \( P_{\text{solution}} \) is the observed vapor pressure of the solution, \( X_{\text{solvent}} \) is the mole fraction of the solvent, and \( P^0_{\text{solvent}} \) is the vapor pressure of the pure solvent.
- Can be used to determine molar mass of a solute.

Raoult’s Law Example

- The vapor pressure of a solution containing 53.6 g glycerin (C\(_3\)H\(_8\)O\(_3\)) in 133.7 g of ethanol (C\(_2\)H\(_5\)OH) is 113 torr at 40°C. Calculate the vapor pressure of pure ethanol at 40°C assuming that glycerin is a nonvolatile, nonelectrolyte solute in ethanol.
Raoult’s Law for Nonideal Solutions

- In a liquid-liquid solution where both components are volatile:
  \[ P_{\text{total}} = P_A + P_B = X_A P_o^A + X_B P_o^B \]
- Liquids obeying Raoult’s Law are called ideal solutions.
- Molecular interactions predict whether or not a solution will obey Raoult’s Law.
- Negative deviation – strong interaction between solvent and solute, \( \Delta H_{\text{soln}} \) is exothermic (\( \text{H}_2\text{O} \) and acetone).
- Positive deviation – weak interactions between solvent and solute, \( \Delta H_{\text{soln}} \) is endothermic (hexane and ethanol).
- No deviation (ideal solution) – solvent-solvent, solute-solute, and solvent-solute interactions are the same, \( \Delta H_{\text{soln}} \) is zero (benzene and toluene).

Nonideal Solutions Example

- A solution is prepared by mixing 0.0300 mol \( \text{CH}_2\text{Cl}_2 \) and 0.0500 mol \( \text{CH}_2\text{Br}_2 \) at 25°C. Assuming the solution is ideal, calculate the composition of the vapor (in terms of mole fractions) at 25°C. At 25°C the vapor pressures of pure \( \text{CH}_2\text{Cl}_2 \) and pure \( \text{CH}_2\text{Br}_2 \) are 133 torr and 11.4 torr, respectively.

Colligative Properties

- Properties of a solution that depend only on the number, and not the identity, of the solute particles.
- Freezing-point depression and boiling-point elevation.
- Colligative properties are useful for determining the molar mass of different compounds.

Boiling-Point Elevation

- A nonvolatile solute will lower the vapor pressure of the solvent.
- A nonvolatile solute elevates the boiling point of the solvent.
- \( \Delta T = K_b m_{\text{solute}} \), where \( K_b \) is a constant characteristic of the solvent called the molal boiling-point elevation constant, \( m \) is the molality of the solute, and \( \Delta T \) is the difference between the boiling point of the pure solvent and the solution.

Example

- A 2.00 g sample of a large biomolecule was dissolved in 15.0 g of carbon tetrachloride. The boiling point of this solution was determined to be 77.85°C. Calculate the molar mass of the biomolecule. For carbon tetrachloride, the boiling-point constant is 5.03°Cxkg/mol, and the boiling point of pure carbon tetrachloride is 76.50°C.

Freezing-Point Depression

- A nonvolatile solute will lower the vapor pressure of the solvent.
- A nonvolatile solute depresses the freezing point of the solvent.
- \( \Delta T = K_f m_{\text{solute}} \), where \( K_f \) is a constant characteristic of the solvent called the molal freezing-point depression constant, \( m \) is the molality of the solute, and \( \Delta T \) is the difference between the freezing point of the pure solvent and the solution.

Example

- The freezing point of t-butanol is 25.50°C and \( K_f \) is 9.1°Cxkg/mol. Usually t-butanol absorbs water on exposure to air. If the freezing point of a 10.0 g sample of t-butanol is 24.59°C, how many grams of water are present in the sample?

Colligative Properties of Electrolyte Solutions

- Electrolytes dissociate into more than one component.
- When dealing with an ionic solution the van’t Hoff factor, \( i \), must be used:.
  \[ i = \text{(moles of particles in solution/ moles of solute dissolved)} \]
- van’t Hoff factor can be less than expected due to ion pairing in solution.