Unit 4 Exam Review: Chapter 10 and Chapter 12

Chapter 9: Electrons in atoms on the periodic table

- **Light:** A form of electromagnetic radiation exhibits both wave like and particle like behavior. Particles of lights are called photons.
- **Electron configuration:** indicates which orbitals are occupied for a particular atom. Orbitals are filled in order of increasing energy and obey the laws of Pauli exclusion principle and Hund's rule.

Chapter 10: Chemical Bonding

- **Lewis Theory:** Chemical bonds are formed when atoms transfer valence electrons (ionic bonding) or share valence electrons (covalent bonding) to attain noble gases electron configurations
- **Molecular Shapes:** The shapes of molecules can be predicted by combining Lewis Theory with Valence Shell Electron Pair Repulsion (VSEPR) theory, where electron groups around the central atom repel one another and determine the geometry of the molecule
- **Electronegativity:** Electronegativity refers to the relative ability of elements to attract electrons within a chemical bond.
- Electronegativity increases as you move to the right across a period in the periodic table and decreases as you move down a column.

Chapter 12: Liquids, Solids, and Intermolecular Forces

- Properties of Liquids
- Properties of Solids
- Manifestations of intermolecular forces: surface tension and viscosity
- Evaporation and Condensation
- Melting, Freezing, and Sublimation

Types of Intermolecular Forces:

- **Dispersion Forces** – Dispersion forces occur between all molecules and atoms due to instantaneous fluctuations in electron charge distribution.
- **Dipole-Dipole Forces** – Dipole-dipole forces exist between molecules that are polar.
- **Hydrogen Bonding** – Hydrogen bonding exists between molecules that have H bonded directly to F, O, or N. Hydrogen bonds are the strongest of the three intermolecular forces.

Types of Crystalline Solids:

- Molecular Solids
- Ionic Solids
- Atomic Solids

Water: Because of its strong hydrogen bonding, water is a liquid at room temperature. Unlike most liquids, water expands when it freezes. Water is highly polar, making it a good solvent for many polar substances. Water is the universal solvent.

Chapter 9: Electrons in Atoms on the Periodic Table

1. Describe the electromagnetic spectrum and draw the entirety of it with corresponding labels.

2. Arrange these three types of electromagnetic radiation – visible light, X-rays, and microwaves – in order of increasing:
   Wavelength:

   Frequency:

   Energy per Photon:

3. Provide an example of the Bohr Model and label all important parts.

4. What is the correct order of the electromagnetic spectrum from the highest to lowest frequency? And for wavelength?

5. How does the rainbow compare to the electromagnetic spectrum in terms of frequency and wavelength?

6. Provide the electron configuration for the following elements on the periodic table:

   Ca:

   Co:

   Cd:

   Zn:

7. How many electrons are unpaired in the orbitals of chlorine? You will probably need to draw the orbital diagram for chlorine using arrows.

8. Write the ground state electron configuration for the magnesium atom.

9. How can you differentiate between core electrons and valence electrons without writing the entirety of an atom's electron configuration?

10. Explain a quantum number (n). What does this number represent?

Chapter 10: Chemical Bonding

1. How do we represent valence electrons?

2. How can the periodic table and electron configuration depict valence electrons?

3. What is the octet rule (duets)?

4. What are resonance structures?

5. What is the difference between electron groups and bonding groups?

Steps for Drawing Lewis Structures:

1. Sum the valence electrons from all atoms.
2. Write the symbols for the atoms to show which atoms are attached to which and connect them with a single bond.
    - Hydrogen is always terminal.
    - The halogens are always terminal.
    - Carbon is often the central atom.
3. Complete the octet of the atoms bonded to the central atom.
4. Place any leftover electrons on the central atom, even if doing so results in more than an octet.
5. If there are not enough atoms to ensure each element has a full octet, try multiple bonds.

1. Discuss the rules of increasing atomic size and list the following elements in order of increasing atomic size: Se, N, Li, Cd, Sc.

2. Describe what occurs when two nonmetals are bonded together.

3. Describe what occurs when one metal and one nonmetal are bonded together.

4. Write the Lewis structure for the following compounds, state their molecular geometry and their bond angles:

CO_2

SO_2

CCl_4

H_2CO

NH_4^+

PCl_3

NO_3^-

H_2O

5. What is electronegativity?

6. Write the Lewis structure for each element: N, C, O, Cl, Ar.

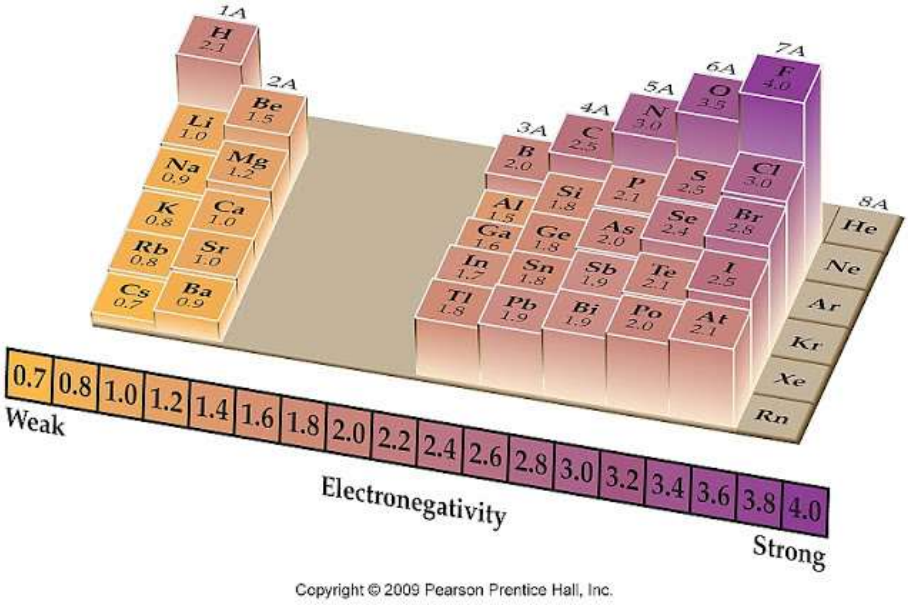
7. Compare bond strength.

8. Why do oil and water not mix?

Electronegativity and Bond Character

Electronegativity	**Bond Character**
2.0 +	
0.4-2.0	
0-0.4	

The Electronegativity Periodic Table



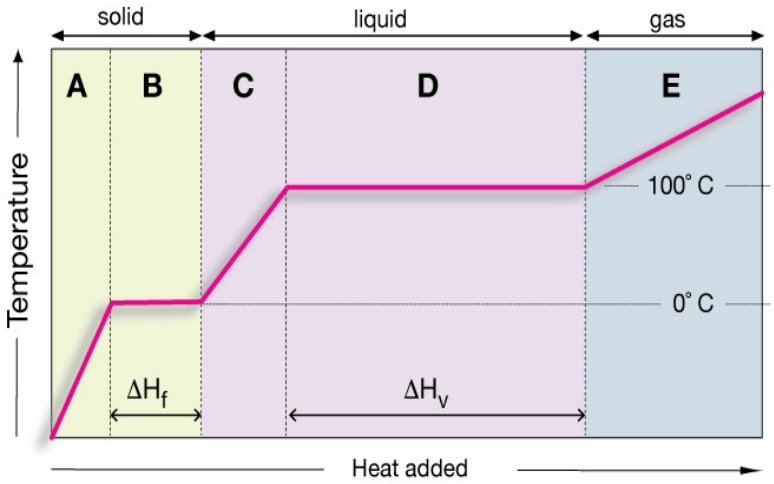
Use the table values to determine what type of bond will form:

1. Sr and O _____
2. Cl and Br _____
3. H and F _____
4. Na and S _____
5. Br and Br _____

Chapter 12: Liquids, Solids, and Intermolecular Forces

1. What properties does water have from strong intermolecular forces?

2. Explain the heating curve of water:



3. What is a dipole moment?

10. How does soap work?

11. Describe the difference in molecular motion between solids, liquids, and gases.

12. Define surface tension, capillary action and viscosity.

13. What factors increase the rate of vaporization?

14. What is vapor pressure and how is it affected by temperature and intermolecular forces?

15. Explain the terms 'endothermic' and 'exothermic' using phase changes.

16. What are the three types of intermolecular forces, and how do they compare in strength?

17. How do you determine if a molecule has dipole–dipole forces?

18. Why does water have a high boiling point despite its low molar mass?

19. Why is water less dense as a solid than as a liquid?

20. If intermolecular forces increase, what happens to the following:

Boiling/Melting Point _____	Capillary _____
Viscosity _____	Vapor Pressure _____
Surface Tension _____	Evaporation Rate _____

21. Define hydrogen bonding. What is an example?

22. What intermolecular interaction is exhibited by all chemicals?

23. List properties and examples of the following intermolecular forces:
    - Ionic
    - Molecular
    - Covalent Network Solids
    - Metallic