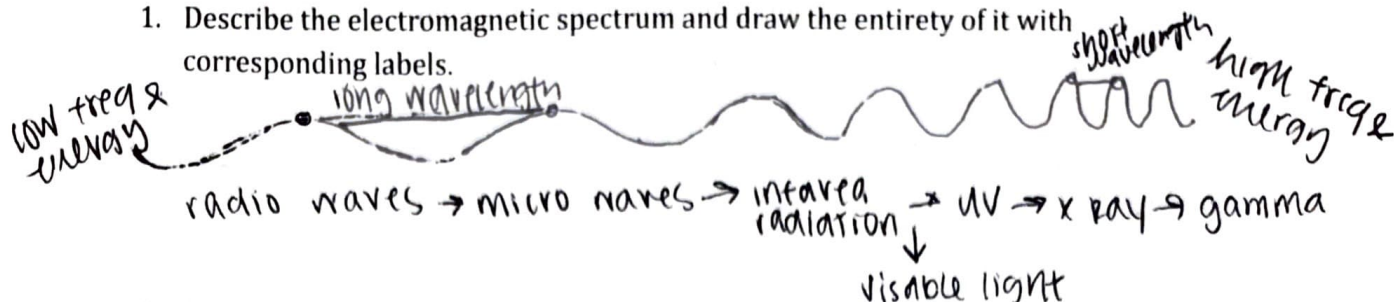


Unit 4 Exam Review: Chapter 9, 10, 12, and 13

Chapter 9: Electrons in Atoms on the Periodic Table

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



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

Wavelength:

X-ray, visible, microwave

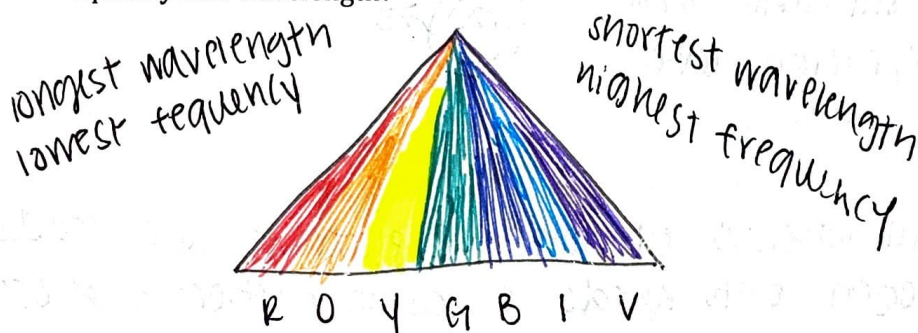
Frequency:

microwave, visible, X ray

Energy per Photon:

↓
same as frequency

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



- Provide the electron configuration for the following elements on the periodic table:

Hunds rule

↓
max # unpaired electrons

Aufbau principle

↓
lower orbitals 1st

Co: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^7$ 2 v.e. * up arrows 1st! fill singly

[Ar] $4s^2 3d^7$

Cd: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10}$ 2 v.e.

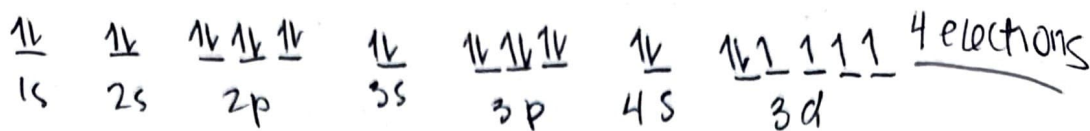
[Kr] $5s^2 4d^{10}$

S: $1s^2 2s^2 2p^6 3s^2 3p^4$ 6 v.e.

$1s^2 2s^2 2p^6 3s^2 3p^4$

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



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



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

look at periodic table columns

I = 1 II = 2 III = 3 IV = 4 V = 5 VI = 6 VII = 7 VIII = 8

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

the subshells and
size/energy level



Chapter 10: Chemical Bonding

1. What is the octet rule (duets)?

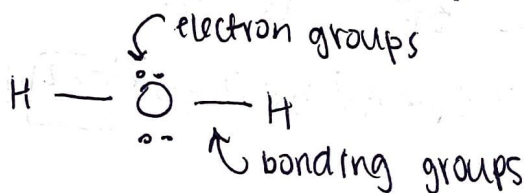
A full valence electron is 8 electrons to be stable.

hydrogen only needs 2 valence electrons to be stable (duet)

2. What are resonance structures?

when a compound can be represented by multiple different structures.

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

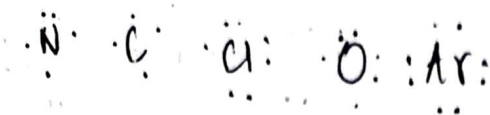


4. Discuss the rules of increasing atomic size and list the following elements in order of increasing atomic size: Se, N, Li, Cd, Sc.

N, Li, Se, Si, Cd.

Unit 4 Exam Review: Chapter 9, 10, 12, and 13

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



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

electrons are shared
covalent, two anions, $(-)(-)$

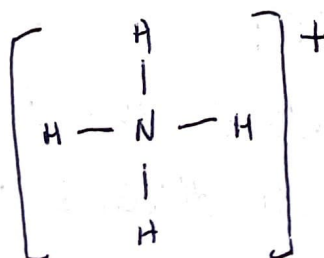
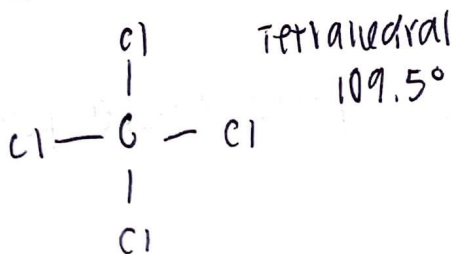
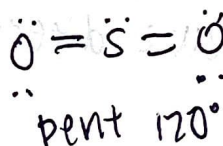
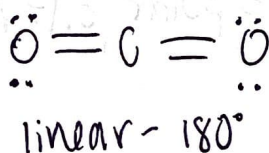
7. Describe what occurs when one metal and one nonmetal are bonded together.

electrons are transferred to nonmetal
cation + anion, ionic, $(+) & (-)$

Steps for Drawing Lewis Structures:

- Sum the valence electrons from all atoms.
- 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.
- Complete the octet of the atoms bonded to the central atom.
- Place any leftover electrons on the central atom, even if doing so results in more than an octet.
- If there are not enough atoms to ensure each element has a full octet, try multiple bonds.

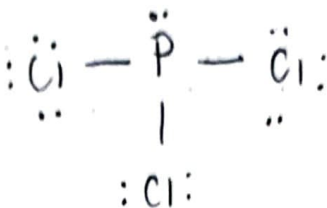
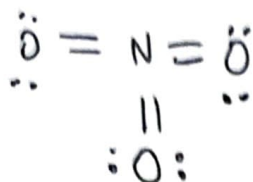
8. Write the Lewis structure for the following compounds, state their molecular geometry and their bond angles:



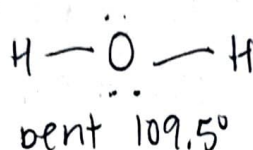
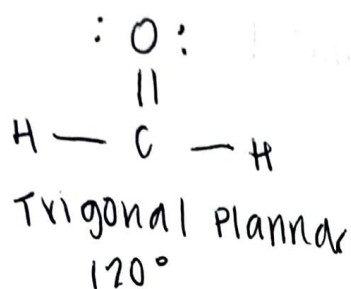
Unit 4 Exam Review: Chapter 9, 10, 12, and 13



Trigonal planar 120°



Trigonal Pyramidal



9. What is electronegativity?

The ability of an atom to attract another atom.

10. Compare bond strength.

Triple & double bonds are shorter and stronger
single bonds are longer and weaker

11. Why do oil and water not mix?

They have different polarities; water = polar, oil = nonpolar

12. What is the universal solvent?

water!

Remember, like dissolves like; polar dissolves polar,
nonpolar dissolves nonpolar.

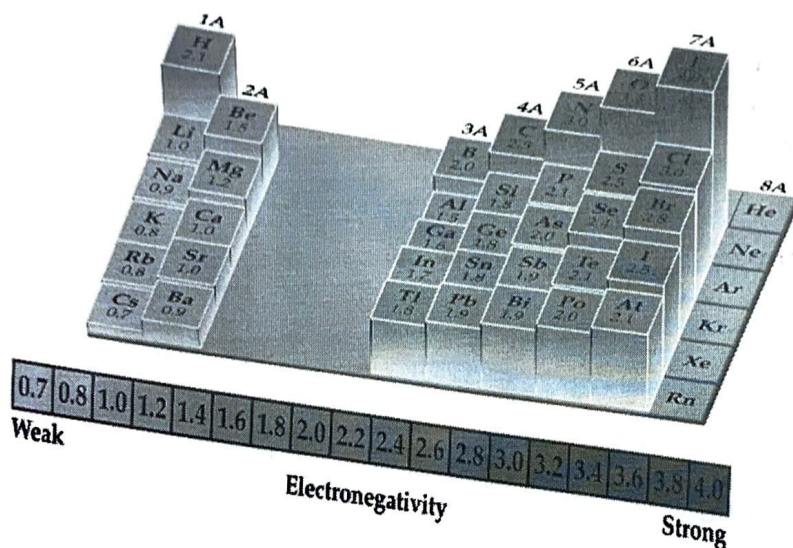
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Electronegativity and Bond Character

Electronegativity	Bond Character
2.0 +	ionic
0.41-1.9	polar covalent
0-0.4	nonpolar covalent

(pure)

The Electronegativity Periodic Table:



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

- Sr and O ionic
- Cl and Br nonpolar
- H and F polar
- Na and S polar
- Br and Br nonpolar

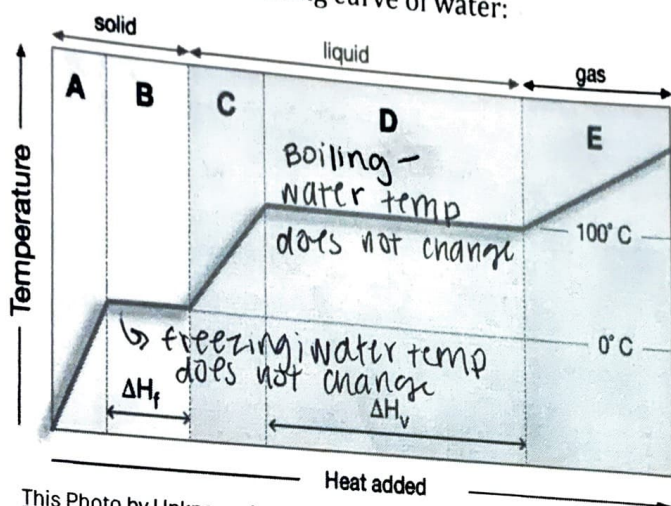
Unit 4 Exam Review: Chapter 9, 10, 12, and 13

Chapter 12 & 13: Liquids, Solids, and Intermolecular Forces

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

high boiling point
high surface tension
low vapor pressure

2. Explain the heating curve of water:



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3. What is a dipole moment?

a measure of the separation of positive and negative charges in a molecule.

10. How does soap work?

it has a polar end and a nonpolar end which allows the water & oil to interact indirectly breaking the grease into smaller units that dissolve water.

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

solids: vibrate

liquids: move freely but close

gases: move rapidly & far apart

Unit 4 Exam Review: Chapter 9, 10, 12, and 13

12. Define surface tension, capillary action and viscosity.

surface tension: energy or work required to increase the surface area of a liquid due to the intermolecular forces

capillary action: the rise of liquids up narrow tubes

viscosity: the resistance of a liquid to flow - increases w/ strong molecular forces, decreases w/ a higher temperature

13. What factors increase the rate of vaporization?

↑ surface area

↑ temperature

↓ intermolecular forces

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

vapor pressure ↑ with temperature

stronger intermolecular forces = lower vapor pressure

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

endothermic: absorbs heat (melting, evaporation)

exothermic: releases heat (freezing, condensation)

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

dispersion, dipole-dipole, hydrogen bonding

weakest

strongest

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

if it is polar and not just dispersion

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

strong hydrogen bonds require more energy to break

20. 19. Define a saturated, unsaturated, and a supersaturated solution.

(mixed them up!) water expands when it freezes due to its molecular structure forming a crystalline lattice.

Unit 4 Exam Review: Chapter 9, 10, 12, and 13

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

saturated = maximum amount of solute. If more solute is added, it won't dissolve.

unsaturated = holding less than maximum amount of solute; dissolves

supersaturated = holding more than the normal max amount of solute; precipitate

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

Boiling/Melting Point $\uparrow$

Capillary $\uparrow$

Viscosity $\uparrow$

Vapor Pressure $\downarrow$

Surface Tension $\uparrow$

Evaporation Rate $\downarrow$

22. Define hydrogen bonding. What is an example?

Attraction between H bonded to N, O, or F

ex) DNA, H_2O , HF

23. What intermolecular interaction is exhibited by all chemicals?

Dispersion

24. List properties and examples of the following intermolecular forces:

o Ionic

high melting point

ex) NaCl , MgO

water soluble

brittle, hard

o Molecular

non polar { low melting point & boiling
non-conducting, insoluble in water

H_2 , CCl_4

polar { higher melting & boiling point
likens soluble

HCl , NH_3

o Covalent Network Solids

Hard, solid

very high melting point

C (diamond)

insoluble in common solvents

SiO_2 (glass, sand, quartz)

o Metallic

variable hardness & melting points

insoluble

Na , Hg , Mg

malleable, ductile