Unit 2- Study Guide
Periodic Trends and Electrons

Name: _____

Periodic Trends

1. What is ionization energy?

The amount of energy needed to lose an e⁻

   a. Explain WHY is changes as you go from Group 1A to Group 8A.

As you move across the PT elements having different amounts of valence electrons

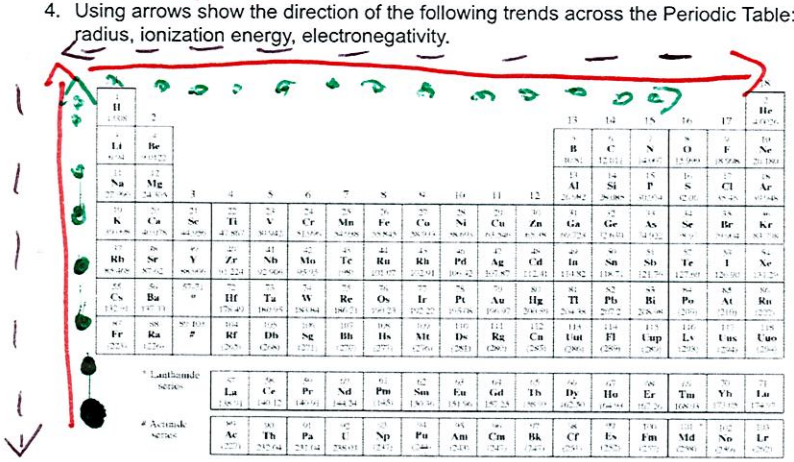
2. Draw figure illustrating how atomic radius is calculated.

1/2 the distance

3. What is electronegativity?

The tendency of an atom to gain an electron

4. Using arrows show the direction of the following trends across the Periodic Table: atomic radius, ionization energy, electronegativity.



Atomic Energy Levels and Electron Configuration

1. Order the following from largest to smallest: ORBITAL, ATOM, LOCATION OF ELECTRON, SUBLEVEL, ENERGY LEVEL

atom, energy level, sublevel, orbital, location

2. $3p^4$

What is the Energy Level? 3

How many electrons are in the sublevel? 4

What type of sublevel is represented? p

3. Complete the following table.

Element	Atomic Number	Orbital illustration *(hint: Arrows)* 1s 2s 2p 2p 2p	Full Electron Configuration
Helium	2	↑↓	$1s^2$
Nitrogen	7	↑↓ ↑↓ ↑ ↑ ↑	$1s^2 2s^2 2p^3$
Neon	10	↑↓ ↑↓ ↑↓ ↑↓ ↑↓	$1s^2 2s^2 2p^6$

4. What is Germanium's atomic number? 32

   a. How many electrons does it have? 32

5. What is noble-gas notation any why is it used to write electron configuration?

It's a way to shorten e⁻ configuration by referencing the last noble gas passed

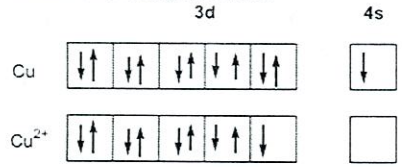
6. Write the electron configuration for ~~Sodium~~ Calcium.

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

   a. Now write the electron configuration using noble-gas notation.

$[Ar] 4s^2$

Use the figure below to answer the following questions.



7. What does each box represent?

an orbital

8. What do the arrows represent?

electrons

9. How many electrons can each orbital hold?

2

10. How many electrons can the d sublevel hold?

10

11. Which is associated with higher energy: a 2s orbital or a 2p orbital?

2p

12. According to the Aufbau principle, which orbital should fill first, a 4s or 3d orbital?

4s

13. How are s orbitals different from p orbitals?

s orbitals have less electrons

14. How many electrons can each of the following orbitals hold?

a. 2s = 2
b. 3p = 6
c. 5f = 14
d. 6d = 10
e. 4p = 6
f. 3d = 10

15. How many "p" orbitals can there be in any energy level? 3

16. What are the ***maximum*** number of electrons in the 3rd principle energy level?

$n=3$

$2n^2 = 2(3^2)$

$=$ 18 electrons

17. How many orbitals are in each of the following sublevels??

a. 4*p* sublevel 3
b. 3*d* sublevel 5
c. 4*f* sublevel 7
d. 2*s* sublevel 1

Quantum Theory

1. How did Bohr expand on Rutherford's model of the atom?

He said that e⁻ have to stay in distinct energy levels

2. Compare the energy of an electron in the ground state and an electron in the excited state.

Electrons have less energy in its grounded state compared to the excited state.

3. When an electron falls from a higher energy level to a lower energy level, how is the energy released?

Energy is released as a photon (light)

4. The further the electron is from the nucleus, the more energy the electron has.

5. A(n) energy level is often thought of as a region of space in which there is a high probability of finding an electron.