AMSAT Chem 1H Mr. Dehne

Name: _____

Date: _____

Per#: ___

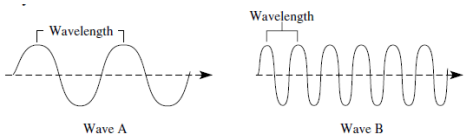
Electron Topic#5

WS#1: The Development of a New Atomic Model

1. The speed of light is ($2.998x10^8$/$3.00x10^9$) meters per second.
2. All waves can be described in terms of their amplitude, wavelength, and (acceleration/frequency).
3. Early in this century, scientists found that light has the characteristics of both waves and (electrons/particles).
4. The (wavelength/frequency) of a wave is the number of complete waves passing a fixed point in a given time.
5. Measuring the distance from a wave crest to the next immediate wave crest gives the (wavelength/frequency) of a wave.
6. The wavelength of microwave radiation is (greater than/less than) the wavelength of visible light.
7. The color of visible light that has the longest wavelength is (red, violet).
8. A heat lamp produces (ultraviolet/infrared) (warm body) radiation.
9. (Ultraviolet/Infrared) radiation causes human skin cells to release melanin.
10. A wave with a high frequency has (long/short) wavelength and (high, low) energy.
11. A wave with a low frequency has (long/short) wavelength and (high, low) energy.
12. A wave with a long wavelength has (high/low) frequency and (high, low) energy.
13. A wave with a short wavelength has (high/low) frequency and (high, low) energy.
14. The brightness of a light depends on the (frequency/amplitude/wavelength) of the light wave.

Short Answer

15. In what way does the photoelectric effect support the particles theory of light?
16. What is the difference between ground state and the excited state of an atom?
17. Under what circumstances can an atom emit a photon?
18. How can energy levels of the atom be determined by measuring the light emitted from an atom?
19. Why does electromagnetic radiation in the ultraviolet region represent a larger energy transition than radiation in the infrared region?
20. Which of the waves shown below has the higher frequency? (The scale is the same for each drawing.) Explain your answer.



21. Fill the following chart.

Decreasing Wavelength ↓	*Type of Electromagnetic Radiation*	*Description of Wave*
	_____ _____ _____ _____	These waves have a long wavelength, low frequency, and low energy.
	_____ _____ _____ _____ _____ _____ _____	These are the colors of the visible spectrum (wavelengths between 750 nm and 400nm).
	_____ _____ _____	These waves have a short wavelength, high frequency, and high energy.

T or F. If false, replace word or words to make it true.

22. Planck proposed that the energy emitted or absorbed by any object is restricted to quanta of particular sizes.
23. We are not aware of quantum effects in the world around us because quanta of energy are very large.

24. In sodium metal, violet light causes the photoelectric effect but red light does not because photons of violet light have less energy than those of red light.
25. Einstein proposed that light consists of photons, which are quanta of energy that behave like particles.
26. In the photoelectric effect, protons are ejected from the surface of a metal when light shines on it.
27. The Planck effect shows that light in the form of a photon can collide with an electron.
28. We are constantly surrounded by low frequency x-rays.
29. Planck's theory relates the frequency of radiation to its energy.
30. The wavelengths of radiation emitted by a hot object shift as its temperature increases.
31. The dual nature of light means that light has the properties of a charge and a wave.

Short Answer

32. Use an analogy to explain the difference between a continuous change and a quantized change.
33. Could you use a highly intense beam of red light to eject electrons from the surface of a metal? Explain your answer.
34. How many different photons of radiation were emitted from excited helium atoms to form the spectrum below? Explain your answer.

Spectrum for helium

On the line at the left, write the letter of the scientist who made the contribution to the quantum theory listed below. Each letter may be used more than once.

a. Planck b. Einstein c. Compton

35. Stated that energy is emitted or absorbed in discrete pieces called quanta.
36. Explained the photoelectric effect in terms of quantized energy.
37. Proved that light consists of tiny particles, or photons.
38. Proposed the idea that light consists of quanta of energy.
39. Related mass and energy through the equation, $E=h\upsilon$.
40. Demonstrated that a photon could collide with an electron.
41. The constant, *h*, which is named after this scientist, is equal to $6.62 \times 10^{-34} \text{J} \cdot \text{s}$.
42. Wrote the equation, $E=mc^2$.
43. How did de Broglie conclude that electrons have a wave nature? Complete

Complete

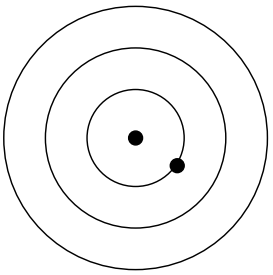
44. Every element has a uniquely characteristic atomic _____spectrum.
45. De Broglie referred to the wavelike behavior of particles as_____.
46. Heisenberg's uncertainty principle states that the position and the _____of a moving object cannot simultaneously be measured and known exactly.
47. _____used Planck's idea of quantization to explain the line spectrum of hydrogen.
48. Bohr labeled each _____in his atomic model by a quantum number.
49. An electron that absorbs a quantum of energy can jump to a level of _____energy, called an excited state of the atom.
50. _____is emitted when an electron jumps from a higher energy level to a lower energy level.
51. The electron microscope makes use of the_____ nature of electrons.

Multiple Choice

52. The Bohr model is an inaccurate model of the atom because there is no way to measure the exact _____of an electron in an atom. a. charge b. path c. mass d. all of the above
53. De Broglie derived a mathematical relationship between the mass and velocity of a moving particle and the _____that it would exhibit. a. motion b. wavelength c. amplitude d. charge
54. When radiation is absorbed by a hydrogen electron, the hydrogen atom changes its ground state to_____.
    a. an excited state b. another atom c. a lower state d. the nucleus of the atom
55. Describe the difference between a continuous spectrum and a line spectrum, and name a source of each kind of spectrum.
56. Why can't you observe the effects of wave motion for a baseball?

Use the diagram below to answer each of the following questions.

57. Label the energy levels $n = 1$, $n = 2$, and $n = 3$ on the Bohr atom.
58. Label the ground state in this atom.
59. Label an excited state in this atom.
60. Draw an arrow to show the direction an electron moves when it absorbs energy.



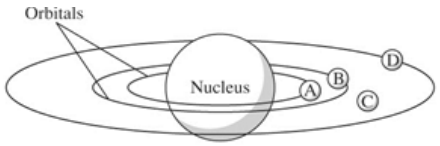
WS#2: The Quantum Model of the Atom / Electron Configurations

electron spin / quantum number / orbital / Pauli exclusion principle/electron density / principal energy levels / quantum-mechanical model

1. A region in space where an electron with a particular energy is likely to be found.
2. The density of an electron cloud.
3. Number designating a principal energy level in an atom.
4. States that each orbital in an atom can hold at most two electrons and that these electrons must have opposite spins.
5. Explains the properties of atoms by treating the electron as a wave and quantizing its energy.
6. The main energy levels in an atom.
7. The clockwise or counterclockwise motion of an electron.

Multiple Choice

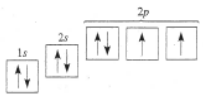
8. The electron cloud is least dense where the probability of finding an electron is
   a. greatest b. lowest c. highly likely d. nonexistent
9. The first principal energy level of the hydrogen atom contains only a(n)
   a. *s* orbital b. *p* orbital c. *d* orbital d. *f* orbital
10. All *p* orbitals are shaped like a. spheres b. doughnuts c. dumbbells d. footballs
11. The 3s orbital differs from the 2s orbital in that it is
    a. smaller b. larger c. a different shape d. more crowded
12. The number of sublevels in each principal energy level equals the
    a. mass of the atom c. quantum number for the energy level
    b. electron density of the atom d. number of electrons in the atom
13. Which sublevels can be found in the fourth principal energy level of an atom?
    a. *s* and *p* b. *s*, *p*, and *d* c. *s*, *p*, *d*, and *f* d. *s*, *p*, *d*, and *g*
14. How does the quantum-mechanical model of the atom describe electrons?
15. Explain the significance of drawing the 90% contour of orbitals.
16. How many quantum numbers are used to describe the properties of electrons in atomic orbitals?
    a. 1 b. 2 c. 3 d. 4
17. A spherical cloud surrounding an atomic nucleus would best represent
    a. an *s* orbital b. an *p* orbital c. an *d* orbital d. an *f* orbital
18. How many electrons can an energy level of $n = 4$ hold?
    a. 32 b. 24 c. 8 d. 6
19. How many electrons can an energy level of $n = 2$ hold?
    a. 32 b. 24 c. 8 d. 6
20. Compared with an electron for which $n = 2$, an electron for which $n = 4$ has more
    a. spin b. particle nature c. energy d. wave nature
21. According to Bohr, which is the point in the figure below where electrons cannot reside?
    a. point A b. point B c. point C d. point D
22. According to the quantum theory, point D in the figure represents
    a. the fixed position of an electron.
    b. the furthest position from the nucleus that an electron can achieve.
    c. a position where an electron probably exists.
    d. a position where an electron cannot exist.



23. Identify each of the four quantum numbers and their properties to which they refer.
24. How did the Heisenberg uncertainty principle contribute to the idea that electrons occupy "clouds," or "orbitals"?
25. Complete the following table.

Principal Quantum Number, ***n***	Number of Sublevels	Types of Orbitals
1		
2		
3		
4		

WS#3: Mixed Review



1. Under what conditions is a photon emitted form an atom?
2. What do quantum numbers describe?
3. What is the relationship between the principal quantum number and the electron configuration?
4. In what way does the figure above illustrate Hund's Rule?
5. In what way does the figure illustrate the Pauli Exclusion Principle?
6. Elements of the fourth and higher main-energy levels do not seem to follow the normal sequence for filling orbitals. Why is this so?
7. How do electrons create the colors in a line-emission spectrum?
8. What is the wavelength of light that has a frequency of 3.000×10^{-4}Hz in a vacuum? How much energy does a photon of this light have? (Ans: 1.0000×10^{12}Hz; $6.626 \times^{-22}$J)
9. What is the energy and frequency of a photon that has a wavelength 600.nm? (Ans: 3.31×10^{-19}J; 5.00×10^{14}Hz)

T or F. If false, replace the word or words to make a true statement.

10. The Pauli Exclusion Principle states that an orbital can hold a maximum of <u>two</u> electrons.
11. The sum of the superscripts in an electron configuration represents the total number of <u>neutrons</u> in an atom.
12. The Aufbau principle states that electrons are added one at a time to the <u>highest</u> energy orbitals available until all the electrons of the atom have been accounted for.
13. An <u>orbital diagram</u> uses arrows to represent the spin of the electrons.
14. The ground state is the <u>least</u> stable energy state of an atom.
15. According to Hund's rule, electrons occupy equal energy orbitals so that a maximum number of <u>unpaired</u> electrons results.

Draw the orbital diagram, ground state and Noble gas electron configurations for each of the following elements.

16. magnesium 17. oxygen 18. aluminum 19. silver 20. scandium 21. indium

Identify the elements with the following electron configurations. Write the chemical symbol for each element.

22. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$
23. $1s^2 2s^2 2p^6 3s^2 3p^2$
24. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^3$
25. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$
26. $1s^2 2s^2 2p^6$
27. Briefly describe how you would use the Aufbau principle, the Pauli Exclusion Principle, and Hund's rule to predict the location of electrons in an atom.

WS#4: Problems WS

Part A: Waves/Energy

1. What is the equation for the speed of light? c = _____. What is the equation for the energy of an electromagnetic wave? E = _____
2. What is the value for the speed of light, c? c = _____. What is the value for Planck's constant, h? h = _____
3. What are the units for c _____, λ _____, h _____, E _____ and υ _____?
4. What do the following symbols represent? υ: (_____), λ: (_____), and c: (_____)
5. Which electromagnetic wave has the longest wavelength? Shortest?
6. Which electromagnetic wave has the highest frequency? Lowest?
7. List the major electromagnetic waves from highest frequency to lowest.
8. Name the colors that make up the visible spectrum.
9. Using E, λ, and υ, devise a relationship between the variables. Use an up arrow to indicate increase and a down arrow for decrease.

Solve for the unknown quantity. Remember all wavelengths must be in meters.

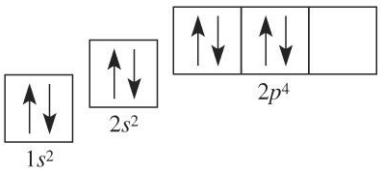
$\upsilon = c/\lambda$ where $c = 3.00 \times 10^8$m/s, $E = \upsilon \bullet h$ where $h = 6.626 \times 10^{-34}$J•s

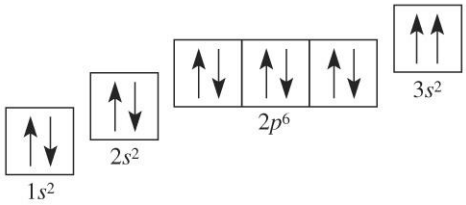
10. If a wave has a wavelength of 523μm, what is the frequency and energy of the wave? (Ans: $\upsilon = 5.74 \times 10^{11}$1/s; $E = 3.80 \times 10^{-22}$J)
11. If a wave has a wavelength of 789nm, what is the frequency and energy of the wave? (Ans: $\upsilon = 3.80 \times 10^{14}$1/s; $E = 2.52 \times 10^{-19}$J)
12. If a wave has a wavelength of 5000Å (1Å = 1×10^{-10}m) , what is the frequency and energy of the wave? (Ans: $\upsilon = 6 \times 10^{14}$1/s; $E = 4 \times 10^{-19}$J)
13. If a wave has a wavelength of 7.4m, what is the frequency and energy of the wave? (Ans: $\upsilon = 4.1 \times 10^7$1/s; $E = 2.69 \times 10^{-26}$J)
14. If a wave has a frequency of 6.6×10^7Hz, what is the wavelength and energy of the wave? (Ans: $\lambda = 4.5$m; $E = 4.4 \times 10^{-26}$J)

15. If a wave has a frequency of 5.14×10^{18}Hz, what is the wavelength and energy of the wave? (Ans: $\lambda = 5.84 \times 10^{-11}$m; $E = 3.41 \times 10^{-15}$J)
16. If a wave has a frequency of 8.97×10^{14}Hz, what is the wavelength and energy of the wave? (Ans: $\lambda = 3.34 \times 10^{-7}$m; $E = 5.94 \times 10^{-19}$J)

Part B: Electron Configurations/Orbital Diagrams/Noble Gas Configurations

1. State the Pauli Exclusion Principle, and use it to explain why electrons in the same orbital must have opposite spin states.
2. Explain the condition under which the following orbital notation for helium is possible: $\underline{\uparrow}_{1s}\ \underline{\uparrow}_{2s}$
3. Which guideline, Hund's Rule or Pauli Exclusion Principle, is violated in the following orbital diagrams?





a. _____ b. _____

4. How many unpaired electrons does each possess?

a. K	____	c. Si	____	e. Ir	____	g. V	____	i. Hg	____
b. Co	____	d. Li	____	f. Bi	____	h. S	____	j. Se	____

5. Identify the element from the following electron configurations:
   a. $1s^2 2s^2 2p^3$ ____
   b. $1s^2 2s^2 2p^6 3s^2 3p^3$ ____
   c. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^1$ ____
   d. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10}$ ____
   e. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^6$ ____
6. For a-h, identify the atomic number of each element (i), draw the orbital diagram (ii), draw the electron configuration and (iii) noble gas (abbreviated) electron configuration (iv), and determine the number of unpaired electrons (e^-) for each atom (v).

a. N	c. V	e. Kr	g. Sn
b. K	d. Zr	f. Br	h. Po

7. Write the noble gas (abbreviated) electron configuration.

a. iridium	b. mercury	c. bismuth	d. radium