WKS 6-8: REVIEW-- Honors
Electron Configurations, Periodic Table, Ions

Name Answer Key
Period ______ ***Date*** ______________

Topics:

- Relationships between energy, frequency & wavelength.
- Calculations using $c = \lambda\nu$ and $E = h\nu$. *(These formulas and the needed constants are given on reference chart!!)*
- Bohr's Model of the hydrogen atom states that electrons have quantized energies. Explain how emission lines give evidence for this model. (Explain by discussing absorption, jumps of electrons, and emission.)
- Color Lab: flame tests, fluorescence, phosphorescence, color of objects as absorption/reflection of visible light
- Wave-particle duality of light and electrons: Particle properties—photoelectric effect and line emission spectra; wave properties; Heisenberg uncertainty principle.
- *Orbitals as e^- probability, not orbits. Rules: Aufbau, Pauli Exclusion Principle and Hund's Rule*
- Electron Configurations: types of orbitals (first energy level seen, number of each), shapes of orbitals (s, p, d), arrow diagrams, writing electron configurations with periodic table
- Determining # of valence electrons, charge of ions in compounds, size of ions vs. size of their neutral atoms
- Labeling families in periodic table (PT without key). What is a period? Group?
- Element Lab type questions: *Anything similar is fair game*. (Mole calcs; density calcs; p, n, e determinations.)

Practice Problems

1) If a beam of red light has a wavelength of 650. nm, how much energy (in J) does one photon of this red light have? What is its frequency?

$$\lambda = 650.\text{ nm} \times \frac{1\times10^{-9}\text{ m}}{1\text{ nm}} = 6.50\times10^{-7}\text{ m};\ c = \lambda\nu \Rightarrow \nu = \frac{c}{\lambda} = \frac{3.00\times10^{8}\text{ m/s}}{6.50\times10^{-7}\text{ m}} = 4.615\times10^{14}\text{ Hz} = \boxed{4.62\times10^{14}\text{ Hz}}$$

$$E = h\nu = \left(6.626\times10^{-34}\text{ J}\cdot\text{s}\right)\left(4.615\times10^{14}\text{ s}^{-1}\right) = 3.058\times10^{-19}\text{ J} = \boxed{3.06\times10^{-19}\text{ J}}$$

Alternately, solve for E first then determine ν from E:

$$E = \frac{hc}{\lambda} = \frac{\left(6.626\times10^{-34}\text{ J}\cdot\text{s}\right)\left(3.00\times10^{8}\text{ m/s}\right)}{6.50\times10^{-7}\text{ m}} = \boxed{3.06\times10^{-19}\text{ J}};\ \nu = \frac{E}{h} = \frac{3.058\times10^{-19}\text{ J}}{6.626\times10^{-34}\text{ J}\cdot\text{s}} = \boxed{4.62\times10^{14}\text{ Hz}}$$

2) A photon of one particular electromagnetic radiation has an energy of 2.94×10^{-17} J. What are the frequency and wavelength of this radiation? What region of the spectrum is it in?

$$E = h\nu \Rightarrow \nu = \frac{E}{h} = \frac{2.94\times10^{-17}\text{ J}}{6.626\times10^{-34}\text{ J}\cdot\text{s}} = 4.437\times10^{16}\text{ Hz} = \boxed{4.44\times10^{16}\text{ Hz}}$$

$$c = \lambda\nu \Rightarrow \lambda = \frac{c}{\nu} = \frac{3.00\times10^{8}\text{ m/s}}{4.437\times10^{16}\text{ /s}} = 6.761\times10^{-9}\text{ m} = \boxed{6.76\times10^{-9}\text{ m}};\ \boxed{\text{X-rays}}$$

Alternately, solve for λ first, then solve for ν:

$$E = \frac{hc}{\lambda} \Rightarrow \lambda = \frac{hc}{E} = \frac{\left(6.626\times10^{-34}\text{ J}\cdot\text{s}\right)\left(3.00\times10^{8}\text{ m/s}\right)}{2.94\times10^{-17}\text{ J}} = \boxed{6.76\times10^{-9}\text{ m}};\ \nu = \frac{c}{\lambda} = \frac{3.00\times10^{3}\text{ m/s}}{6.761\times10^{-9}\text{ m}} = \boxed{4.44\times10^{16}\text{ Hz}}$$

Answers: 1) 4.62×10^{14} Hz; 3.06×10^{-19} J; 2) 4.44×10^{16} Hz; 6.76×10^{-9} m; X-rays

3) True or False? Green light has lower energy than orange light. False

4) True or False? If an electron falls a relatively short distance, light of relatively low energy would be emitted. True

5) What does it mean for electron energy levels to be *quantized*?
Quantization means that electrons can occupy only certain discrete energy levels.

6) Why does an excited electron naturally fall towards the nucleus?
coulombic attraction
The negative electrons are attracted to the positive nucleus by the electrostatic force.

7) When electricity is sent through an emission tube filled with hydrogen gas, distinct energies of light are emitted. Thus, a bright line or emission spectrum is obtained. Explain how this emission spectrum gives experimental evidence for the concept that the electron in hydrogen can exist in only distinct energy levels.
- The emission spectrum of hydrogen consists of distinct lines. Thus, distinct energies of light are emitted.
- Distinct energies can be emitted only when the electron loses distinct amounts of energy.
- Thus, the electron must be limited in what "jumps down" it is allowed to make.
- Thus, the electron must be limited in what energies it can have. (The energies of the electron are quantized.)
- Saying this in another way: The electron in a hydrogen atom can only exist in distinct energy levels.

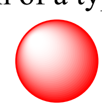
8) Fluorescent substances only fluoresce (emit visible light) when ultraviolet (UV) light shines on the object.

9) When you see an object as red, it is because …
a) all colors are being absorbed b) only red is absorbed c) all colors except red are absorbed

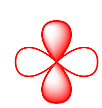
10) Explain why a black shirt gets hotter in the sun than a white shirt. (What is absorbed? What is emitted? What happens to e^-'s?)
- A black shirt absorbs all visible colors of light (reflects none). Thus, electrons are excited into higher energy levels. When electrons "fall" back down to lower energy levels, heat can be emitted. Thus, the shirt gets hotter.
Infrared light
- A white shirt absorbs no visible colors of light (reflects all). Thus, the electrons cannot be excited, so no heat is emitted.
- NOTE: Heat is also released by the sun. Thus, a white shirt does get somewhat hotter because heat is directly hitting it.

11) What does an orbital (or electron cloud) represent? How is an orbital different from an orbit?
- An orbital is a region where there is high probability of finding the electron that does not travel in a predictable, defined path. An orbit has a defined path by which the electron travels.
- There is no way to predict with certainty where the electron will be in the future. One can only know the probability of finding it in a particular region.

12) Draw a representation of a typical "s" orbital, "p" orbital and a "d" orbital. How many orientations are there of each type?


s orbital
1 orientation


p orbital
3 orientations
p_x, p_y, p_z


d orbital
5 orientations

13) Describe the photoelectric effect. How does it support the conclusion that light can have properties of particles?
- When light is shone on metal, electrons will only be ejected if the light has the minimum energy required for that metal. (For example, one metal might require yellow light—another might require uv light.)
- If one increases the intensity of the light, this only increases the number of photons, so this increases the number of electrons emitted. However, if none of the photons have the minimum energy required, no electrons are emitted.
- In order for an electron to be emitted, the needed light energy must hit at the exact location of the electron. Thus the light must have "packets of energy" (with at least the needed minimum energy) that hit each electron. This means that light is acting like a particle (has a distinct location). We call these particles of light, **photons.**

14) What does the Heisenberg uncertainty principle state? How do electrons in atoms satisfy the uncertainty principle?

- Heisenberg uncertainty principle states that one cannot simultaneously know both the location and energy of anything with certainty.
- All electrons in atoms must satisfy the uncertainty principle because they have properties of waves. For an electron in an orbital, one knows the energy precisely, but the location is somewhat uncertain. (Only know the *probable* location of e−)

15) How many electrons can fit in the third main energy level?____18_____What is another name for a row?_period_____

16) What sublevel should become filled after the 5s sublevel is filled?_4d______

17) What is the maximum number of electrons that can fit in any one orbital?_2________

18) Questions about Germanium:

a) Fill out the arrow diagram for Ge to the right

b) How many unpaired electrons does Ge have?_2_____

c) How many valence electrons does Ge have?_4______

d) Why do the electrons remain unpaired in the 4p sublevel?

- Electrons are all negative, so they repel each other.
- Thus, if available, electrons will spread out into another orbital of the same energy. (Hund's Rule)

5s ____
4p $\uparrow$ __ $\uparrow$ __ ___
3d $\uparrow\downarrow$ $\uparrow\downarrow$ $\uparrow\downarrow$ $\uparrow\downarrow$ $\uparrow\downarrow$
4s $\uparrow\downarrow$
3p $\uparrow\downarrow$ $\uparrow\downarrow$ $\uparrow\downarrow$
3s $\uparrow\downarrow$
2p $\uparrow\downarrow$ $\uparrow\downarrow$ $\uparrow\downarrow$
2s $\uparrow\downarrow$
1s $\uparrow\downarrow$

19) Write the electron configurations for these elements. (Use noble gas notation.)

a) Te $[Kr]5s^2 4d^{10} 5p^4$

b) W $[Xe]6s^2 4f^{14} 5d^4$

20) What is the name of the family of elements in Group 1/IA_Alkali Metals______ Group 2/IIA?_Alkaline Earth Metals_________ Group 17/VIIA?_Halogens______ Group 18/VIIIA?_Noble Gases_

21) Answer these questions concerning the element whose electrons configuration ends with $4p^4$.

a) The element is _Se_____. Write out its electron configuration $[Ar]4s^2 3d^{10} 4p^4$

b) How many valence electrons does it have?____6______

c) How many unpaired electrons does it have?_2_________

d) When in compounds, what is its most common ion?__Se^{2-}__________

e) What is the electron configuration of this most common ion?_$[Ar]4s^2 3d^{10} 4p^6$_

f) Is this ion smaller or larger than its neutral atom?_larger__

g) If a particular isotope of this element had 40 neutrons, what would be its mass number?_34 + 40 = 74_____________

h) What type of element is this-- metal, nonmetal or metalloid? _nonmetal_________

i) What would be the mass of a 0.650 mole sample of this element? $0.650\ \text{mol} \times \frac{78.97\ \text{g}}{1\ \text{mol}} = \boxed{51.3\ \text{g}}$