Practice Problems: *Wavelength, Frequency, Energy content of One Quantum of Light.*

Examples:

I. A certain photon of light has a wavelength of 422 nm. What is the frequency of the light?

\[ \lambda = \frac{c}{v} \]
\[ \frac{3.0 \times 10^8 \text{ m/s}}{422 \times 10^{-9} \text{ m}} = 7.11 \times 10^{14} \text{ Hz} \]

II. What is the energy of a quantum of light from part I.

\[ E = h \nu \]
\[ = 6.626 \times 10^{-34} \text{ J} \cdot \text{s} \cdot (7.11 \times 10^{14} \text{ Hz}) \]
\[ = 4.71 \times 10^{-19} \text{ J} \]

1. What is the energy of a quantum of light with a frequency of \( 7.39 \times 10^{14} \text{ Hz} \)?

\[ E = 6.626 \times 10^{-34} \text{ J} \cdot \text{s} \cdot (7.39 \times 10^{14} \text{ Hz}) \]
\[ = 4.90 \times 10^{-19} \text{ J} \]

2. What is the wavelength of the quantum of light in question 1?

\[ \lambda = \frac{c}{\nu} \]
\[ 3.0 \times 10^8 \text{ m/s} = \frac{4.06 \times 10^{-7} \text{ m}}{7.39 \times 10^{14} \text{ Hz}} \]
\[ \approx 4.06 \times 10^{-7} \text{ m or } 406 \text{ nm} \]

3. A certain red light has a wavelength of 680 nm. What is the frequency of the light?

\[ \lambda = 3.0 \times 10^8 \text{ m/s} = 680 \times 10^{-9} \text{ m} \]
\[ \nu = \frac{4.41 \times 10^{14} \text{ Hz}}{680 \times 10^{-9} \text{ m}} \]
4. What is the energy of a quantum of light from question 3?

\[ E = \frac{6.626 \times 10^{-34} \text{ J s}}{4.41 \times 10^4 \text{ Hz}} \]

\[ = 2.92 \times 10^{-19} \text{ J} \]

5. A certain blue light has a frequency of 6.91 x 10^{14} \text{ Hz}. What is the wavelength of the light?

\[ 3.0 \times 10^8 \text{ m/s} = \lambda \left(6.91 \times 10^{14} \text{ Hz}\right) \]

\[ 4.34 \times 10^{-7} \text{ m} = 434 \text{ nm} \]

6. What is the energy of a quantum of light from question 5?

\[ E = \frac{6.626 \times 10^{-34} \text{ J s}}{6.91 \times 10^{14} \text{ Hz}} \]

\[ = 4.58 \times 10^{-19} \text{ J} \]

7. The energy for a quantum of light is 2.84 x 10^{-19} \text{ J}. What is the wavelength of this light?

\[ 2.84 \times 10^{-19} \text{ J} = \frac{6.626 \times 10^{-34} \text{ J s}}{v} \]

\[ v = 4.29 \times 10^{14} \text{ Hz} \]

\[ 3.0 \times 10^8 \text{ m/s} = \lambda \left(4.29 \times 10^{14} \text{ Hz}\right) \]

\[ \lambda = 6.99 \times 10^{-7} \text{ m} = 699 \text{ nm} \]
Information: Energy of Sublevels

Each sublevel has a different amount of energy. For example, orbitals in the 3p sublevel have more energy than orbitals in the 2p sublevel. The following is a list of the sublevels from lowest to highest energy:

1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s, 5f, 6d...

To help you, here is the above list with the orbitals included. Recall that each blank represents an orbital:

1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f

Note that d and f sublevels appear to be out of place. This is because they have extra high energies. For example, the 3d sublevel has a higher energy than a 4s sublevel and the 4f sublevel has a higher energy than the 6s sublevel.

When electrons occupy orbitals, they try to have the lowest amount of energy possible. (This is called the Aufbau Principle.) An electron will enter a 2s orbital only after the 1s sublevel is filled up and an electron will enter a 3d orbital only after the 4s sublevel is filled. Recall that only two electrons can fit in each orbital. (This is called the Pauli Exclusion Principle.) When two electrons occupy the same orbital, they must spin in opposite directions—one clockwise and the other counterclockwise.

Critical Thinking Questions

1. a) How many electrons would an atom need to have before it can begin filling the 3s sublevel?
   
   10

   b) What is the first element that has enough electrons to have one in the 3s sublevel? (Use your periodic table.)

   Na

2. a) How many electrons would an atom need to have before it can begin filling the 3d sublevel?

   20

   b) What is the first element that has enough electrons to begin placing electrons in the 3d sublevel?

   Sc
**Information:** Hund’s Rule

Electrons can be “paired” or “unpaired”. Paired electrons share an orbital with their spins parallel. Unpaired electrons are by themselves. For example, boron has one unpaired electron. Boron’s orbital diagram is below:

```
1s   2s   2p
      ↑   ↑   ↑
   Boron
```

If we added one more electron to boron’s orbital diagram we will get carbon’s orbital diagram. One important question is: where does the next electron go? The electron has a choice between two equal orbitals—which of the 2p orbitals will it go in?

**Choice A** or **Choice B**

```
1s   2s   2p
      ↑   ↑   ↑
```

Hund’s rule tells us which of the above choices is correct. Hund’s rule states: when electrons have a choice of entering two equal orbitals they enter the orbitals so that a maximum number of unpaired electrons result. Also, the electrons will have parallel spins. Therefore Choice A is carbon’s actual orbital diagram because the p electrons are in separate orbitals and they have parallel spins.

The following are the electron orbital diagrams for the next elements, nitrogen and oxygen. Notice that nitrogen’s 2p electrons are all unpaired to obey Hund’s Rule. The 2p electrons are forced to begin pairing up in oxygen’s configuration.

```
1s   2s   2p
      ↑   ↑   ↑
   Nitrogen
```

```
1s   2s   2p
      ↑   ↑   ↑
   Oxygen
```

**Critical Thinking Questions**

3. How many “unpaired” electrons are in a nitrogen atom?

4. Why does carbon’s sixth electron have to go into another p orbital? Why can’t it go into a 2s orbital? Why can’t it go into a 3s orbital?

5. Write the electron orbital diagram for phosphorus.

```
1s    2s    2p    3s   3p  3d
    ↑    ↑    ↑   ↑   ↑   ↑
```

6. Write the electron orbital diagram for arsenic.

```
1s    2s    2p    3s   3p   4s  3d
    ↑    ↑    ↑   ↑   ↑   ↑   ↑
```
7. Compare the orbital diagrams for nitrogen (see information section above), phosphorus (question 5) and arsenic (question 6). What is similar about the electrons in the last sublevel for each of them?

Each have 3 unpaired e⁻ in the last p sublevel.

**Information: Electron Configurations vs. Orbital Diagrams**

The electron orbital diagram of an atom can be abbreviated by using what is called electron configurations. The following is the electron configuration for carbon: 1ˢ² 2ˢ² 2p². The following is the electron configuration for several elements whose orbital diagrams are given above:

- Carbon: 1ˢ² 2ˢ² 2p²
- Nitrogen: 1ˢ³ 2ˢ² 2p³
- Oxygen: 1ˢ² 2ˢ² 2p⁴

**Critical Thinking Questions**

8. What are the small superscripts (for example, the ⁴ in oxygen) representing in an electron configuration?

The number of electrons.

9. What information is lost when using electron configurations instead of orbital diagrams? When might it be more helpful to have an orbital diagram instead of an electron configuration?

Whether e⁻ are paired or unpaired.

10. How many unpaired electrons are in a sulfur atom? What did you need to answer this question—an orbital diagram or an electron configuration?

[Diagram: 1s² 2s² 2p⁴ 3s² 3p⁴ 1] 12.

11. Write the electron configuration for zirconium (atomic # = 40).

1s² 2s² 2p⁶ 3s² 3p⁶ 3d¹⁰ 4s² 3d¹⁰ 4p⁶ 5s² 4d²

12. Write the configuration for argon (atomic # = 18).

1s² 2s² 2p⁶ 3s² 3p⁶

13. Write the electron configuration for calcium (atomic # = 20). Notice that calcium has all of argon's electrons plus two additional ones in a 4s orbital.

1s² 2s² 2p⁶ 3s² 3p⁶ 4s²
Electron configuration practice

1. Draw a horizontal orbital diagram for Sulfur (S)
2. Draw a vertical orbital diagram for Sodium (Na) and Copper (Cu)
3. Write FULL electron configurations for C, K, Rb, Co, Ar, Ti, Pb
4. Write shorthand (noble gas configurations) for N, Ti, Se, Mn, Sr, Hg, Te, Re, Fr, Po, Ce, Lu
5. Be aware of exceptions. Write out electron configurations (full or shorthand) for Cr, Mo, Cu, Ag
6. Write out the FULL electron configuration for N\(^3\), Mg\(^{2+}\), Zn\(^{2+}\), Ag\(^+\), Cl\(^-\), K\(^+\), O\(^{2-}\)

1. \[ 3s^2 \quad 3p^1 \quad \underline{1s^2 2s^2 2p^3} \]

2. \[ \text{Na:} \quad 1s^2 2s^2 2p^6 3s^1 \quad \text{Cu:} \quad 1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10} 4p^2 \]

3. \[ \text{C:} \quad 1s^2 2s^2 2p^2 \quad \text{K:} \quad 1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 \]

4. \[ \text{Rb:} \quad 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5 5s^1 \]

5. \[ \text{Co:} \quad 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^7 \quad \text{Ar:} \quad 1s^2 2s^2 2p^6 3s^2 3p^6 \]

6. \[ \text{Ti:} \quad 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} \quad \text{Pb:} \quad 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5 5s^2 4d^{10} 5p^{10} 6p^2 \]

7. \[ \text{Sr:} \quad [Kr] 5s^2 \quad \text{Hg:} \quad [Xe] 6s^2 4f^{14} 5d^{10} \]

8. \[ \text{Te:} \quad [Kr] 5s^2 4d^{10} 5p^4 \]

9. \[ \text{Re:} \quad [Xe] 6s^2 4f^{14} 5d^5 \]

10. \[ \text{Fr:} \quad [Rn] 7s^1 \quad \text{Po:} \quad [Xe] 6s^2 4f^{14} 5d^{10} 6p^4 \]

11. \[ \text{Ce:} \quad [Xe] 6s^2 4f^{14} 5d^1 \]

12. \[ \text{Lu:} \quad [Xe] 6s^2 4f^{14} 5d^1 \]

13. \[ \text{Cu:} \quad [Ar] 4s^1 3d^{10} \quad \text{Ag:} \quad [Kr] 5s^1 4d^{10} \]

14. \[ \text{N:} \quad [He] 2s^2 2p^3 \]

15. \[ \text{Sr:} \quad [Kr] 5s^2 \]

16. \[ \text{N}_3^2^-: 1s^2 2s^2 2p^6 \]

17. \[ \text{Mg}^{2+}: 1s^2 2s^2 2p^6 \]

18. \[ \text{Zn}^{2+}: 1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} \quad \text{(loses } e^- \text{ in } 4s) \]

19. \[ \text{Ag}^+: 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^1 5s^1 4d^{10} \quad \text{(lose } e^- \text{ in } 5s) \]

20. \[ \text{Cl}^-: 1s^2 2s^2 2p^6 3s^2 3p^6 \]

21. \[ \text{K}^{+}: 1s^2 2s^2 2p^6 3s^2 3p^6 \]

22. \[ \text{O}^{2-}: 1s^2 2s^2 2p^6 \]
Electromagnetic Radiation and Quantum Theory Questions

What you will be given for the test:

- A periodic table (you cannot use the blank one where you filled in the electron configuration)
- $c=\lambda v$
- $E=nh$
- $c=3.0\times10^8 \text{ m/s}$
- $h=6.626\times10^{-34} \text{ Js}$

For the test, you should be able to answer these questions....

1. What does ROY G. BV represent? 10/26
2. How does the energy of a red photon compare to that of a blue photon? Explain. 10/26
3. According to the Bohr model of the hydrogen atom, how does the hydrogen atom emit light? 10/26
4. What is the difference between an electron at ground state and an excited electron? 10/26
5. What accounts for different color lines (red, blue-green, blue, and violet) in the emission spectrum of the hydrogen atom? 10/26
6. Can two different elements produce the same identical emissions spectrum? 10/26
7. What is a "quantum" of energy? 10/26
8. What kind of relationship do frequency and wavelength have? Energy and frequency? 10/25
9. What speed does all light travel at? 10/25
10. Be able to solve energy, frequency, and wavelength problems. 10/26
12. What is an energy level? 10/26
13. What is a sublevel? 10/26
14. What is an orbital? 10/26
15. Describe the orbitals in the 4 sublevels. How do they differ from one another? 10/26
16. When filling atomic orbitals with electrons, describe the three principles you should follow. 10/26
17. How does a 1s orbital compare to a 2s orbital? How are they similar? How do they differ? 10/25
18. What type of orbital starts each and every new energy level (n)? 10/26
19. Identify the types of atomic orbitals found in the 4th energy level (n=4)? How many of each type are present? How many total electrons can go in the 4th energy level? 10/26
20. What is the ending electron configuration for each group (1A-8A) on the periodic table? 10/26
21. Determine the total number of electrons that can occupy 1 s orbital, 3 p orbitals, 5 d orbitals, 7 f orbitals 10/25
22. How many electrons does it take to completely fill the third energy level? 10/26
23. Identify the following elements given their ENDING electron configuration
   a. $1s^2$  
   b. $3s^3p^5$  
   c. $4s^3d^8$  
24. Draw the vertical orbital diagram for the element phosphorus (P). 10/26
25. Draw the horizontal orbital diagram for calcium (Ca). 10/26
26. Be able to identify an element based on its electron configuration. 10/26
27. Write the full electron configuration for the following elements: B, V, Fe, Br, Sn 10/26
28. Write the noble gas configurations for Y, At, I, Ba 10/26
29. Write the electron configuration for $S^{2-}$, $P^3^+$, $K^+$, $Si^{4+}$ 10/26
30. What is a valence electron? Know that elements in the same group have the same number of valence electrons. 10/26
31. Be able to tell how many valence electrons are in an atom based on the electron configuration. 10/26
32. Draw Lewis structures for all main group elements in the 3rd and 4th period. 10/26
33. In terms of electron configuration/orbital filling, explain why an atom will not have more than eight valence electrons. 10/26
34. Choose one element on the periodic table with an exception and write its electron configuration. 10/26
35. Give one example of a paramatic element and diamatic element using electron configurations.