Chapter 5 Homework:
Periodicity and Atomic Structure

1) An infrared wave is measured to have a wavelength of $5.6 \times 10^4$ nm. What is its frequency?

$$5.4 \times 10^{12} \text{ s}^{-1}$$

2) Almost all commercially available microwave ovens employ radiation with a frequency of $2.45 \times 10^9 \text{ s}^{-1}$. Calculate the wavelength of this radiation.

$$0.122 \text{ m}$$

3) An FM radio station broadcasts at a frequency of $9.87 \times 10^7 \text{ s}^{-1}$ (98.7 MHz). Calculate the wavelength of the radio waves. What is the energy of one photon of this radiation? ($h = 6.626 \times 10^{-34} \text{ J} \cdot \text{s}$)

$$\lambda = 3.04 \text{ m}; E = 6.54 \times 10^{-26} \text{ J}$$

4) Light with a wavelength of 400 nm strikes the surface of metallic cesium in a photocell and the energy of the electrons ejected from the metal is $1.54 \times 10^{-19} \text{ J}$. Verify Planck’s constant using this information. (What is the color of light used on the metal?)

$$400 \text{ nm} \rightarrow 7.50 \times 10^{14} \text{ s}^{-1}; h = 6.63 \times 10^{-34} \text{ J} \cdot \text{s}$$

5) Both blue and green light eject electrons from the surface of potassium. In which case will the ejected electrons have the higher average kinetic energy? Explain your answer.

Blue: with a wavelength of ~400, this equates to an energy of $4.97 \times 10^{-19} \text{ J}$
Green light: wavelength of ~550 corresponds to an energy of $3.61 \times 10^{-19} \text{ J}$

6) a. Describe the difference between absorption and emission spectra. How is each one obtained?

Absorption spectra are obtained when electrons in an element absorb energy from a source (e.g., light shining on the sample). Emission spectra are obtained when electrons in an element are excited to a higher energy level and eventually drop back down to the ground state, giving off energy as they do so. Those energy jumps correspond to specific wavelengths of light (i.e., specific colors).

7) Why do elements exhibit line spectra instead of continuous emission spectra?

As electrons jump back down to the ground state, they emit energy of a specific value. This amount of energy corresponds to a wavelength of light and a certain color. Since each element only has specific
energy levels available to it, only specific energy transitions are allowed. Therefore, emission spectra for a given element only exhibit specific colors of light.

8) What is meant by the word orbital?

Region of space with a high probability of finding an electron.

9) Describe the difference between a shell, a subshell, and an orbital.

Shells: rows of the periodic table. Defined by numbers \( \geq 1 \).
Subshells: groups of orbitals that make up a shell (e.g., s, p, d, and f)
Orbital: groups of regions that make up a subshell (s subshell has 1 orbital, p subshell has 3 orbitals, d subshell has 5 orbitals)

10) What happens to the probability of finding an electron as you move further away from the nucleus?

Probability decreases. Electrons will most likely be found near the nucleus and less likely to exist on the outer edge of an orbital.

11) Describe and draw the shapes of the \( s \), \( p \), and \( d \) orbitals. How many orientations exist for each set?

\( s \): spherical
\( p \): dumb-bell shaped (two lobes)
\( d \): intersecting dumb-bell (four lobes)

12) a. What is the maximum number of electrons that can exist in any orbital?
   2

b. How many electrons can exist in the \( p \) subshell?
   6

c. How many electrons can exist in the \( 2p_x \) orbital?
   2

d. How many \( p \) orbitals are in the \( n = 2 \) shell?
   3

e. How many electrons can exist in the \( d \) subshell?
   10

f. How many subshells are allowed in the \( n = 3 \) shell?
   3

g. How many \( d \) orbitals are allowed in the \( n = 3 \) shell?
   5

h. How many electrons can exist in the \( n = 3 \) shell?
13) Describe the Aufbau Principle, the Pauli Exclusion Principle, and Hund’s Rule. Give an example of each.
Aufbau: electrons are placed in orbitals of increasing energy; Pauli: no two electrons can have the same four quantum numbers; Hund’s: electrons will be placed in orbitals of a subshell with one in each orbital and then paired up after each orbital is half-filled.

14) Write electron configurations for the following atoms (using the long-hand notation):
   a) Ca: 1s\(^2\) 2s\(^2\) 2p\(^6\) 3s\(^2\) 3p\(^6\) 4s\(^2\)
   b) Cl: 1s\(^2\) 2s\(^2\) 2p\(^6\) 3s\(^2\) 3p\(^5\)
   c) Ga: 1s\(^2\) 2s\(^2\) 2p\(^6\) 3s\(^2\) 3p\(^6\) 4s\(^2\) 3d\(^{10}\) 4p\(^1\)
   d) Ti: 1s\(^2\) 2s\(^2\) 2p\(^6\) 3s\(^2\) 3p\(^6\) 4s\(^2\) 3d\(^2\)

15) Write electron configurations for the following atoms or ions (using short-hand notation):
   a) Br: [Ar] 4s\(^2\) 3d\(^{10}\) 4p\(^5\)
   b) Cr: [Ar] 4s\(^1\) 3d\(^5\)
   c) O: [He] 2s\(^2\) 2p\(^4\)
   d) Ag: [Kr] 5s\(^1\) 4d\(^{10}\)
   e) Cu: [Ar] 4s\(^1\) 3d\(^{10}\)
   f) Rb: [Kr] 5s\(^1\)
   g) Mo: [Kr] 5s\(^1\) 4d\(^5\)

16) Draw orbital diagrams (valence electrons only) for the following atoms and ions:
   a) Cr: \(\uparrow\uparrow\uparrow\uparrow\uparrow\uparrow\uparrow\) \(\uparrow\uparrow\uparrow\uparrow\) \(\downarrow\downarrow\downarrow\downarrow\downarrow\downarrow\downarrow\)  
      4s \ 3d
   b) O: \(\uparrow\downarrow\) \(\uparrow\downarrow\uparrow\uparrow\) \(\uparrow\) \(\downarrow\) \(\downarrow\)  
      2s \ 2p
   d) Ag: \(\uparrow\uparrow\uparrow\uparrow\uparrow\uparrow\uparrow\)  
      \(\downarrow\downarrow\downarrow\downarrow\downarrow\downarrow\downarrow\)
17) Why do elements with the same electron configurations exhibit similar chemical properties (e.g., chemical reactivity)?

Elements with the same electron configurations will tend to gain or lose the same number of electrons in a chemical reaction and therefore will undergo the same types of reactions.