

## Chapter 7 Focus Questions

### Section 1

1. List and define three characteristics of electromagnetic waves. Draw a picture of a wave and label wavelength and amplitude.
2. What is the relationship between wavelength and frequency?
3. What is the equation that relates wavelength and frequency?
4. What is the speed of light?

### Section 2

1. Before 1900, matter and waves seemed very distinct from each other. Who proved this statement incorrect and what did his results support?
2. Define quantization.
3. What is the value of Planck's constant?
4. Is the energy of matter continuous?
5. Since energy is transferred in "packets" or "quanta," it demonstrates that energy also behaves like a particle (matter). Einstein took this step further and said that Electromagnetic radiation (although it is a wave) is also a stream of particles called "photons." What is the equation for the energy of a photon?
6. What does the energy of a photon depend on? (Look at the variables in the equation).
7. What is the equation that states that energy has mass?
8. Calculate the mass of a photon of an x-ray ( $\lambda=1 \times 10^{-10} \text{m}$ ).
9. What is the "dual nature of light?"
10. What did deBroglie say about matter?
11. What is deBroglie's equation?
12. What does the wavelength of matter depend on?
13. Calculate your wavelength traveling at 1m/s. You must convert your weight from pounds to kilograms.
14. Compare that value to an electron traveling at  $1 \times 10^7 \text{m/s}$ . See Ex. 7.3.
15. What is diffraction?
16. What is a diffraction pattern?
17. Does this prove wave or particle nature of light?
18. Explain Davisson and Germer's experiment.
19. What did their experiment verify?
20. Do photons primarily exhibit wave or particle like behavior?
21. Do electrons primarily exhibit wave or particle like behavior?
22. Do large objects primarily exhibit wave or particle like behavior?

### Section 3

1. How is the emission spectrum of hydrogen produced?
2. What is the difference between a continuous spectrum and a line spectrum?
3. What is the significance of the line spectrum of hydrogen?
4. What is the equation to calculate the energy of a photon emitted when an electron jumps between energy levels?

5. Why isn't the emission spectrum continuous?

#### **Section 4**

1. What was Bohr's great development?
2. What was the idea that led to Bohr's model?
3. Why don't electrons travel in a straight line?
4. Do electrons eventually collapse into the nucleus?
5. Draw a picture of Bohr's model.
6. What is the equation that gives the energy levels available to an electron in a hydrogen atom? What does  $Z$  represent? What does  $n$  represent? What does the negative sign represent?
7. Are electrons more tightly bound to the nucleus closer or farther from the nucleus?
8. Does a very negative number mean tightly or loosely bound to the nucleus?
9. If an atom lost energy to make a transition, is the transition more or less stable?
10. What is the equation for the wavelength of the photon emitted when an electron jumps to a ground state (or absorbed when an electron jumps to an excited state)?
11. What is equation 7.2 used for?

#### **Section 5**

1. Heisenberg, deBroglie, and Schrödinger developed what approach to the atomic model?
2. What part of an electron did Schrödinger emphasize?
3. Electrons bound to the nucleus were thought to produce a standing wave. What is a standing wave?
4. What is the requirement for a standing wave to occur?
5. Draw a standing wave and label the nodes.
6. In the example of a standing on a string that is bound to both ends, there are nodes at each end. As a result, how many wavelengths are allowed?
7. Draw the next standing wave pattern (5 nodes) for figure 7.10. How many half wavelengths are there?
8. How do you determine the size the circumference of an energy level will have?
9. In Schrödinger's equation, what does the  $H$  stand for?  $\Psi$ ?  $E$ ?
10. The total energy of an atom is equal to the sum of PE and KE. Where does the PE come from and the KE come from?
11. Is there just one solution for a wave function to Schrödinger's equation?
12. Each wave function that Schrödinger's equation solves for has an associated energy. What are these wave functions called?
13. What is the lowest energy wave function called?
14. What is the Heisenberg Uncertainty Principle (in words and mathematically)?
15. What does  $\Delta x$  stand for?  $\Delta mv$ ?
16. What is the value of the minimum uncertainty for the product of  $\Delta x \bullet \Delta(mv)$ ?

17. As the uncertainty of  $\Delta x$  increases (in other words, the more you don't know about an object's position), what happens to the uncertainty of the objects momentum?
18. What is the square of a wave function?
19. What does the square of the wave function for the 1s orbital give?
20. Where is the probability of finding the electron in the hydrogen atom the greatest?
21. At what distance is the electron most likely to be found in the H atom?
22. Is it always found at this distance?
23. What is the exact size of a 1s orbital?

## **Section 6**

1. What are quantum numbers?
2. What is the quantum number,  $n$ ?
3. As  $n$  increases, what does  $E$  do?
4. As  $n$  increases, what does the size of the orbital do?
5. Is the energy less or more negative as  $n$  increases? Is the electron, therefore, more or less tightly bound to the nucleus?
6. What is the quantum number,  $l$ ?
7. If  $n=3$ , then  $l=0,1,2$ ? What letters do these represent? If  $n=4$ , what are the possible values of  $l$ ? What letters do these represent?
8. What is the quantum number,  $m_l$ ?
9. If  $n=3$ , then  $l=0,1,2$ , so that  $m_l$  must have values of  $-2,-1,0,1,2$ . Let's look at  $l=1$ : If  $l=1$ ,  $m_l=-1,0,1$ . Since  $l=1$  corresponds to the p sublevel, there must be three different orientations of orbitals, because there are three values for  $m_l$  (recall from general chemistry that the p sublevel has 3 orbitals-along the x, y, and z axes.)
10. For the d sublevel ( $l=2$ ),  $m_l$  can be  $-2,-1,0,1,2$ . How many orbitals does the d sublevel contain?
11. For the f sublevel ( $l=3$ ), what are the possible values of  $m_l$ ? How many orbitals does the f sublevel contain?
12.  $n=1$  contains how many sublevels? What is it (letter)?
13.  $n=2$  contains how many sublevels? What are they (letters)?
14.  $n=3$  contains how many sublevels? What are they (letters)?
15.  $n=4$  contains how many sublevels? What are they (letters)?
16. How many orbitals does s have?
17. How many orbitals does p have?
18. How many orbitals does d have?
19. How many orbitals does f have?
20. What quantum number signifies the principle energy level?
21. What quantum number signifies the type of sublevel (s,p,d,f)?
22. What quantum number signifies the orientation of each orbital?
23. For the s sublevel, what does  $l$  equal?
24. For the p sublevel, what does  $l$  equal?
25. For the d sublevel, what does  $l$  equal?
26. For the f sublevel, what does  $l$  equal?

## **Section 7**

1. What is a nodal surface?
2. Draw a picture of a 3s orbital and label the nodal surfaces.
3. How many nodal surfaces do you think the 4s orbital has?
4. What happens to the shape of s as n increases in number? In other words, as the energy level increases, what happens to each s orbital?
5. As the energy level increases, what happens to each p orbital?
6. What is the basic shape of a p orbital?
7. Name the p orbitals with x, y, and z
8. What n number does p start?
9. What n number does d start?
10. Name the d orbitals with x, y, and z
11. As the energy level increases, what happens to each d orbital?
12. Are f orbitals formed in bonding?
13. As the energy level increases, what happens to each f orbital?
14. What does it mean to degenerate?
15. How does the energy of the 4s, 4p, 4d, and 4f orbitals compare to the hydrogen atoms?
16. Note: for multielectron atoms, 4s, 4p, 4d, and 4f orbitals are not degenerate. However, the three p orbitals there are, as are the 5 d orbitals, as are the 7 f orbitals
17. What is the ground state for the hydrogen atom?

## **Section 8**

1. Why was it necessary to assume that electrons must spin?
2. How does a charge produce a magnetic moment?
3. How many spin states are there?
4. What values can the magnetic spin quantum number,  $m_s$  have?
5. What is that Pauli Exclusion Principle?
6. According to the Pauli Exclusion Principle, what quantum numbers may have the same for two electrons? Which quantum number must be different?
7. In an atom, suppose there is an electron that occupies the  $3d_{xz}$  orbital. Give all four quantum numbers for this electron (assume the electron is spinning "up").
8. In an atom, there is another electron that occupies the  $3d_{xz}$  orbital. Give all four quantum numbers for that electron.

## **Section 9**

1. What is a polyelectronic atom?
2. What are the three energy contributions in an atom?
3. State the electron correlation problem and what approximation must be made?
4. Describe electron shielding.
5. Although polyelectronic atoms have orbitals similar in shape to the H atom, there is something that is very different. What is that difference?
6. What has greater energy, a 2s orbital or a 2d orbital?

7. What is the penetration effect?
8. As the electron gets farther away from the nucleus, does the energy of the electron increase or decrease?

### **Section 10**

1. Why was the periodic table originally developed?
2. Who is generally given credit to the present form of the periodic table and why?
3. When did Mendeleev predict Ge? What year was it actually discovered? Were all of the properties pretty accurately predicted by Mendeleev?
4. In row 8/Group 3 of Mendeleev's table (Fig. 7.23), what is Di known as today? (Look at the modern periodic table).
5. In row 6/Group 7, what element did Mendeleev predict that has atomic mass of about 100? (Look at the modern periodic table).
6. What is the row # and group # where Mendeleev predicted Ge?
7. Since the undiscovered element with atomic number 113 is similar in properties to thallium, what group must this undiscovered element belong to? How about number 114?
8. What is the difference between Mendeleev's periodic table and the periodic table we use today?

### **Section 11**

1. In your own words, what is the Aufbau Principle?
2. In your own words, what is Hund's rule?
3. Draw the orbital diagram for chlorine and state how many unpaired electrons there are.
4. Write the electron configuration for chlorine (long form and the short form using the noble gas).
5. What are valence electrons?
6. Why are valence electrons so important?
7. What electron orbitals are considered valence electrons?
8. Why do elements in the same group have similar properties?
9. What is the electron configuration for Br?
10. What is the expected electron configuration for Cr? What is the actual configuration?
11. Since half filled sublevels tend to be favored, write the electron configuration for Mo?
12. Why do the s sublevels fill before the d sublevels?
13. Why would an electron occupy a 5d orbital before a 4 orbital at times?
14. How many valence electrons does group 1 have? Group 2? Group 3? Group 4? Group 5? Group 6? Group 7? Group 8?
15. Why is it often difficult to predict the electron configuration of the transition metals, lanthanides, and actinides?

## **Section 12**

1. What is ionization energy?
2. What is the first ionization energy? The second? The third?
3. Why is there such a large jump in the third and fourth ionization energies for aluminum?
4. For a group 2 element, where would the large jump occur?
5. Going across the period, I.E. (increases, decreases). Why?
6. Going down a group, I.E. (increases, decreases). Why?
7. Draw the orbital diagrams for Be and B, then explain why the I.E. is lower for B than for Be.
8. Draw the orbital diagrams for N and O, then explain why the I.E. is lower for O than for N?
9. Draw the orbital diagrams for P and S, then explain why the I.E. is lower for S than for P.
10. What is electron affinity?
11. For atoms that form negative charges, is the electron affinity endothermic or exothermic? Why does that make sense?
12. What is the trend as you go from left to right? Why?
13. Use orbital diagrams to explain why  $C^-$  is more stable than  $N^-$  (and therefore why  $C^-$  has a more negative value of E.A.).
14. Explain why  $O^-$  is more stable than  $N^-$ .
15. What is the trend for E.A. as you go down a group? Why?
16. Why does F have a smaller value of E.A. than expected from the trend?
17. Why are atomic radii often called covalent atomic radii?
18. When atoms form bonds, what happens to their electron clouds?
19. What is the trend in atomic radius as you go left to right across a period? Why?
20. What is the trend in atomic radius as you go down a group? Why?

## **Section 13**

1. What determines an atom's chemistry?
2. What are the special names of groups 1, 2, 7, and 8?
3. Where are the metals found? What charges do they form? Why do they have lower I.E. than nonmetals?
4. Why would atoms with low ionization energies be very reactive?
5. Where are the nonmetals found? What charges do they form? Why do they have higher I.E. than metals?
6. Why are the upper right hand elements more reactive (except for the noble gases)?
7. Why are the noble gases unreactive?
8. Where are semimetals located?
9. Why is H a nonmetal?
10. Why does density generally increase down a group?
11. Is there a general trend among all groups in melting and boiling points?
12. Are metals oxidizing or reducing agents? Nonmetals?
13. Why is Cs a better reducing agent than Li?

14. Predict which would be the better reducing agent, Sr or Mg?
15. Why is hydration energy?
16. Why are polar water molecules more attracted to  $\text{Li}^+$  than  $\text{Na}^+$  or  $\text{K}^+$ ?
17. Since water makes Li an ion before Na, or K, what is the trend for the greater reducing agent in water?
18. If water ionizes Li very well, why do Na and K react more vigorously with water rather than with Li?
19. Is the energy change for a reaction related to the reaction rate?