Atomic Structure: Chapter Problems

Bohr Model

Class Work

1. Describe the nuclear model of the atom.
2. Explain the problems with the nuclear model of the atom.
3. According to Niels Bohr, what does “n” stand for?
4. Using wavelengths emitted from a hydrogen atom, a student finds that the value of “n” in the Balmer Series of hydrogen is 4.5. Is this value accurate? Justify your answer.
5. According to Bohr, why do atoms emit light?
6. How do electrons get to an excited state?
7. Which of the following transitions would produce the greatest amount of energy: $2 \rightarrow 1$, $3 \rightarrow 2$, or $4 \rightarrow 3$?
8. A photon with a frequency of $4.5 \times 10^{15}$ Hz is emitted when an electron returns to its ground state. What is the energy difference between the ground state and the excited state of this electron?
9. A photon with a wavelength of 720 nm is emitted when an electron returns to a lower energy state. What is the energy difference between the orbits of this electron?
10. What is the energy of the light emitted by an atom if the excited electron absorbs a photon with a frequency of $7.0 \times 10^{15}$ Hz?
11. What is the wavelength of a photon emitted from an atom if its energy is $5.65 \times 10^{-19}$ J?
12. Electromagnetic radiation with a wavelength of 531 nm appears as green light to the human eye. The energy of one photon of this light is $3.74 \times 10^{-19}$ J. Thus, a laser that emits $2.3 \times 10^{-2}$ J of energy in a pulse of light at this wavelength produces _______ photons in each pulse.
13. How can emission spectra be used to identify unknown elements?

Homework

14. Explain how atomic structure violates Coulomb’s law.
15. What is an emission spectrum?
16. How did emission spectra of gases affect our understanding of atomic structure?
17. Describe the Bohr model of an atom.
18. Define “ground state”.
19. Define “excited state”.
20. Which of the following transitions would produce the greatest amount of energy: $1 \rightarrow 2$, $3 \rightarrow 1$, or $2 \rightarrow 1$?
21. A photon with a frequency of $7.75 \times 10^{14}$ Hz is emitted when an electron returns to its ground state. What is the energy difference between the ground state and the excited state of this electron?

22. A photon with a wavelength of 450 nm is emitted when an electron returns to a lower energy state. What is the energy difference between the orbits of this electron?

23. What is the energy of the light emitted by an atom if the excited electron absorbs a photon with a wavelength of 622 nm?

24. What is the frequency of a photon emitted from an atom if its energy is $4.38 \times 10^{-18}$ J?

25. A green spectral line (525 nm) is emitted from an atom. What is the frequency and energy of this photon?

26. Why did scientists need to improve upon the Bohr model?

**Quantum Mechanics**

**Class Work**

27. *What implication does the equation $\lambda = \frac{h}{mv}$ have on how we view matter or anything with momentum?

28. *What is the de Broglie wavelength of a 12.0 gram bullet traveling at the speed of sound? The speed of sound is 331 m/s.

29. *What is the wavelength of an electron that has a velocity of $3.5 \times 10^7$ m/s?

30. *What is the frequency of an electron that has a velocity of $2.12 \times 10^8$ m/s?

31. *What is the energy of an electron that has a velocity of $5.8 \times 10^7$ m/s?

32. Explain how the double slit experiment has impacted our understanding of atomic structure.

33. What is the Heisenberg Uncertainty Principle?

34. How does quantum mechanics differ from Newtonian mechanics?

**Homework**

35. *Why is it not possible for an electron to continue in a set orbit around the nucleus like a planet around the sun?

36. *What is the de Broglie wavelength of a 10.0 gram whip traveling at the speed of sound? The speed of sound is 331 m/s.

37. *What is the wavelength of an electron which has a velocity of $1.2 \times 10^8$ m/s?

38. *What is the frequency of an electron which has a velocity of $9.0 \times 10^7$ m/s?

39. *What is the energy of an electron which has a velocity of $6.334 \times 10^8$ m/s?

40. Define wave function.

41. Why does the dual nature of matter make it difficult to observe very small particles like electrons?

42. State Schrödinger’s equation.

**The Quantum Model**

**Class Work**

43. Name and describe the 4 quantum numbers.
44. Give the number of orientations for each type of orbital.
45. What orbital shapes can be found in the n = 1 level?
46. What is the maximum number of electrons that could be in the n = 3 level?
47. **The spin quantum number has how many possible values?
48. **What are the possible sets of quantum numbers for an electron in a 1s orbital?
49. **What are the possible sets of quantum numbers for an electron in a 2p orbital?
50. How does “n” in the Bohr model differ from “n” in the quantum model?

**Homework**
51. Name the shape of orbitals in the first two subshells.
52. Give the maximum number of electrons in each subshell.
53. What orbital can be found in the n = 3 level?
54. What is the maximum number of electrons that could be in the n = 2 level?
55. **When n = 4, how many possible values exist for m_l?
56. **What is the equation for angular quantum numbers in terms of n?
57. **What are the possible sets of quantum numbers for an electron in a 3d orbital?
58. Which subshell has the higher energy: 3d or 4s?

**Electron Configurations**

**Class Work**
59. Draw the energy level diagram for Iron.
60. Draw the energy level diagram for Sulfur.
61. Draw the energy level diagram for Argon.
62. Draw the energy level diagram for Neon.
63. How does the Aufbau Principle affect the way you draw an energy level diagram?
64. Full shells make an atom more stable. Using what you know about electrons and energy levels, do you think that oxygen will gain or lose electrons to achieve a full shell? Explain your answer using an energy level diagrams.
65. What is the electron configuration of Iron?
66. What is the electron configuration of Bromine?
67. What is the electron configuration of Lithium?
68. Identify the following element: 1s^22s^22p^63s^23p^2
69. Can an electron configuration be 1s^32s^1? Explain.
70. *What is the electron configuration for excited state lithium?

**Homework**
71. Draw the energy level diagram for Titanium.
72. Draw the energy level diagram for Strontium.
73. Draw the energy level diagram for Krypton.
74. Draw the energy level diagram for Copper.
75. How does Hund’s Rule affect how you draw an energy level diagram?
76. How does the Pauli Exclusion Principle affect how you draw an energy level diagram?
77. Using what you know about electrons and energy levels, do you think that Calcium will gain or lose electrons to achieve a full shell? Explain your answer using an energy level diagram.
78. What is the electron configuration of Strontium?
79. What is the electron configuration of Nickel?
80. What is the electron configuration of Francium?
81. *What is the electron configuration for F^-?
82. Identify the following element: \(1s^22s^22p^63s^23p^24s^23d^{10}4p^65s^1\)
83. Why is it incorrect to write the electron configuration of selenium, \(1s^22s^22p^63s^23d^{10}3p^24s^24p^4\)
84. *What is the electron configuration for excited state potassium?

**Free Response**

1. The currently accepted best model of the atom is the quantum mechanical model. Trace the evolution of this atomic model by describing each of the following and the problems that were resolved in the next model.
   a. Plum Pudding Model
   b. Nuclear Model
   c. Bohr Model
   d. Quantum Model

2. The emission spectra of three elements is given below. Use these spectra to answer the following questions.

![Spectra Diagram]

   a. Explain why these emission spectra occur.
   b. What is the significance of the hydrogen spectrum?
c. Sodium has two spectral lines in the visible range: 588.9950 nm and 589.5924 nm. How much energy is produced by each of these lines?
d. Predict the energy levels of mercury’s spectral lines compared to those of sodium.
e. Draw the energy level diagram for hydrogen and sodium.

3. The emission spectra of two elements is given below. Use these spectra to answer the following questions.

![Emission Spectra](Helium.png)

![Emission Spectra](Neon.png)

Frequency

a. How would the Bohr model explain the greater number of spectral lines in neon vs. helium?
b. Two of helium’s spectral lines have wavelengths measuring 501.567 nm and 587.562 nm. What color are these lines?
c. What is the wavelength and energy of helium’s spectral line with a frequency of 6.710 x 10^{14} Hz?
d. Predict the energy levels of neon’s spectral lines compared to those of helium.
e. Write the electron configuration for these elements.

4. Carbon-14 is an unstable isotope of carbon with a half-life of 5730 years.
   a. Draw the energy level diagram for this element. Explain how you used Hund’s Rule, the Aufbau principle, and the Pauli Exclusion Principle to construct your diagram.
   b. Carbon-13 is a stable isotope of carbon. How will its energy level diagram compare to that of Carbon-14?
   c. Carbon-14 undergoes beta decay. Will the energy level diagram of the product be different? Explain.
   d. Write the electron configuration for carbon-14.
e. *State the possible quantum numbers for the lowest energy electrons.

5. $^{60}_{27}Co^* \rightarrow ^{60}_{27}Co + \gamma$
   a. Identify the type of reaction above.
   b. Write the electron configurations for $^{60}_{27}Co^*$ and $^{60}_{27}Co$.
   c. Will the energy level diagram of these two species be identical? Justify your response.
   d. *State the possible quantum numbers for the highest energy electrons of $^{60}_{27}Co$. 
**Answers**

1. In the nuclear model of the atom, protons (and neutrons) are housed in a small, dense nucleus. Electrons surround the nucleus in an area of mostly empty space.

2. If electrons are electrically attracted to nucleus and would, therefore, have centripetal acceleration in order to orbit the nucleus. Accelerating charges emit radiation, so the electron would lose energy and therefore velocity and crash into the nucleus.

3. According to Bohr, “n” stands for the orbit or energy level of the electron.

4. No, this is not accurate. “n” can only be whole numbers; there are no half-levels in the Bohr model of the atom.

5. According to the Bohr model, atoms emit light because excited electrons are returning to lower energy states, emitting the energy difference. This energy always has a specific wavelength because the electrons can only exist in set orbits.

6. Electrons get to an excited state by absorbing the energy of a photon.

7. \(2 \rightarrow 1\)

8. \(3.0 \times 10^{-18} \text{ J}\)

9. \(2.76 \times 10^{-19} \text{ J}\)

10. \(4.6 \times 10^{-18} \text{ J}\)

11. 352 nm

12. \(6.2 \times 10^{16} \text{ photons}\)

13. Because each element has a different number of protons and electrons, they’re “n” values differ from each other. Each element thus produces its own emission spectra which allows scientists to identify elements.

14. Coulomb’s law says that oppositely charged objects attract. Electrons should “fall into” the positively charged nucleus, but this does not happen, somehow violating the law.

15. An emission spectrum is the frequencies of light emitted from an atom.

16. Emission spectra showed that electrons only emit radiation at certain wavelengths and frequencies, and, therefore, energy levels. This indicating that electrons could be found in specific energy levels or orbits.

17. In the Bohr model of the atom, the nucleus is at the center of the atom with electrons orbiting it at set distances, similar to the way planets orbit the Sun.

18. The ground state is the lowest energy state that an electron can reside.

19. Excited states are any energy state above the ground state.

20. \(3 \rightarrow 1\)

21. \(5.14 \times 10^{-19} \text{ J}\)

22. \(4.41 \times 10^{-19} \text{ J}\)

23. \(3.19 \times 10^{-19} \text{ J}\)

24. \(6.61 \times 10^{15} \text{ Hz}\)

25. \(5.71 \times 10^{14} \text{ Hz}; 3.78 \times 10^{-19} \text{ J}\)

26. The Bohr model was only able to explain the emission spectrum of hydrogen.

27. Anything with momentum has a wavelength and is therefore a wave. Small masses, like electrons have a larger wavelength than larger masses.

28. \(2 \times 10^{-34} \text{ m}\)

29. \(2.1 \times 10^{-11} \text{ m}\)

30. \(6.18 \times 10^{19} \text{ Hz}\)

31. \(1.6 \times 10^{-14} \text{ J}\)
32. The double slit experiment illustrates that electrons are both a particle and a wave. It also shows that we cannot determine the path of an individual electron but we can determine the probable distribution of electrons.

33. The position and momentum of an electron cannot be known simultaneously.

34. Quantum mechanics differs from Newtonian mechanics because it takes into account the wave properties of matter. Newtonian mechanics could be used to predict the exact motion of an object. Quantum mechanics can be used to determine the probable motion of an object.

35. Because the electrons would need to continually radiate energy because they would have to be accelerating all the time in order to keep them from colliding with the positive nucleus.

36. \(2 \times 10^{-34}\) m

37. \(6.1 \times 10^{-12}\) m

38. \(1.1 \times 10^{19}\) Hz

39. \(3.66 \times 10^{-13}\) J

40. The wave function describes the state and behavior of an electron. It is tells us the probability of finding an electron at a single point.

41. Observing these particles changes their momentum, so position and momentum cannot be simultaneously known.

42. \(H\Psi = E\Psi\)

43. Principal quantum number \((n)\) – energy level of the electron; angular quantum number \((l)\) – shape of the orbital; magnetic quantum number \((m)\) – orientation of the orbital; spin quantum number \((m_s)\) – direction of the electron spin

44. S orbitals – 1 orientation; p orbitals – 3 orientations; d orbitals – 5 orientations; f orbitals – 7 orientations

45. S only

46. 18

47. 2, +1/2 and -1/2

48. 1 0 0 +½; 1 0 0 -½

49. 2 1 1 +½; 2 1 1 -½; 2 1 0 +½; 2 1 0 -½; 2 1 -1 +½; 2 1 -1 -½;

50. The “n” in the Bohr model is a single level. The “n” in the quantum model is broken down into sublevels.

51. S – spherical; p – lobe shaped

52. S – 2; p – 6; d – 10; f – 14

53. S, p, and d

54. 8

55. 7

56. \(l = n-1\)

57. 3 2 2 +½; 3 2 2 -½; 3 2 1 +½; 3 2 1 -½; 3 2 0 +½; 3 2 0 -½; 3 2 -2 +½; 3 2 -2 -½; 3 2 -1 +½; 3 2 -1 -½

58. 3d
62.

63. The Aufbau principle says you must fill in the 1s orbital before moving to the 2s orbital before moving to the 2p orbitals and so on.

64. Oxygen will gain 2 electrons rather than lose the four it has at its highest energy level.

65. 1s\(^2\) 2s\(^2\) 2p\(^6\) 3s\(^2\) 3p\(^4\) 4s\(^2\) 3d\(^6\)

66. 1s\(^2\) 2s\(^2\) 2p\(^6\) 3s\(^2\) 3p\(^6\) 4s\(^2\) 3d\(^{10}\) 4p\(^5\)

67. 1s\(^2\) 2s\(^1\)

68. Silicon

69. No, it violates the Pauli Exclusion Principle

70. 1s\(^2\) 2p\(^1\)

71.
73. 

74. 

75. Hund's Rule says you can't start pairing electrons until all the orbitals on an energy level are full.
76. The Pauli Exclusion Principle says you must draw one arrow up and the other down to represent electrons because they have opposite spins.
77. Calcium will lose 2 electrons. It has a full 3s subshell but to get a full shell it would need to lose those 2 or gain another 16 to have a full 3 shell. Losing the 2 is easier.
78. $1s^2 \ 2s^2 \ 2p^6 \ 3s^2 \ 3p^6 \ 4s^2 \ 3d^{10} \ 4p^6 \ 5s^2$  
79. $1s^2 \ 2s^2 \ 2p^6 \ 3s^2 \ 3p^6 \ 4s^2 \ 3d^8$  
80. $1s^2 \ 2s^2 \ 2p^6 \ 3s^2 \ 3p^6 \ 4s^2 \ 3d^{10} \ 4p^6 \ 5s^2 \ 4d^{10} \ 5p^6 \ 6s^2 \ 4f^{14} \ 5d^{10} \ 6p^6 \ 7s^1$  
81. $1s^2 \ 2s^2 \ 2p^6$  
82. Rubidium  
83. Electron configurations are written in terms of increasing energy levels, 3d has a higher energy level than 4s  
84. $1s^2 \ 2s^2 \ 2p^6 \ 3s^2 \ 3p^6 \ 4s^2 \ 3d^{10} \ 4p^6 \ 5s^2 \ 4d^{10} \ 5p^6 \ 6s^2 \ 4f^{14} \ 5d^{10} \ 6p^6 \ 7s^1$  

Free Response  
1. Evolution of atomic theory  
   a. The plum pudding model of the atom consisted of a positively charge sphere with electrons embedded in it. It was developed because it was support by the finding that protons were much more massive than electrons and by Coulomb's law of attraction between oppositely charged particles. It was disproved by the gold foil experiment which demonstrated that the atom was mostly empty space with a small dense core.  
   b. The nuclear model consisted of an atom of mostly empty space with a dense core containing the protons and neutrons. It could not explain the lack of continuous spectrum of energy being emitted from the atom.  
   c. The Bohr model explained the quantized spectra lines as set orbits where the electrons could exist without emitting energy, but it could not explain why the electrons did not emit energy in these orbits.  
   d. The quantum model is based on the wave-particle duality of matter and can only be used to determine the probable path and location of electrons but not their exact orbit or location. It cannot explain why or how an electron can behave as both a particle and a wave.  
2. Three emission spectra  
   a. According to the Bohr model, emission spectra occur in quantized amounts because electrons orbit the nucleus at set distances and when they fall from higher to lower energy orbits they emit light at a set wavelength.  
   b. The hydrogen spectrum was the first and only spectrum that could be adequately explained by the Bohr model.  
   c. $3.373 \times 10^{-19} \text{ J}; 3.369 \times 10^{-19} \text{ J}$  
   d. Mercury's spectral lines would produce more energy because they are at a higher frequency. $E = h\nu$.  

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3. Two emission spectra
   a. Because Neon has more electrons it is expected to have more energy levels/orbits.
   b. 501.567 nm – blue; 587.562 nm – green/yellow
   c. 446.8 nm; 4.446 x 10^{14} J
   d. Neon’s spectral lines have lower energy than the majority of helium’s visible spectrum lines.
   e. 1s^2; 1s^22s^22p^6

4. Carbon-14
   a. Pauli Exclusion principle only allows you to put two electrons in each orbital, drawing one arrow up and the other down. The Aufbau principle says to fill the 1s before the 2s, and then move to 2p. Hund’s Rule says to put one arrow (electron) in each 2p space before doubling up.
   b. It doesn’t Carbon-13 and Carbon-14 have the same number of electrons in their neutral state.
   c. Carbon-14 undergoing beta decay produces nitrogen-14. Nitrogen has 7 electrons in its neutral state so the energy level diagram would be different – an arrow (electron) would be added to the empty 2p orbital.
   d. 1s^22s^22p^2
   e. 1 0 0 +½; 1 0 0 -½

5. Cobalt
   a. Gamma decay
   b. 1s^22s^22p^63s^23p^64s^23d^64p^1; 1s^22s^22p^63s^23p^64s^23d^7
c. No, they will have the same number of electrons but because one of the electrons in $^{60}_{27}Co^*$ is excited and violates the Aufbau principle.

d. $3 \ 2 \ 2 \ +\frac{1}{2}; \ 3 \ 2 \ 2 \ -\frac{1}{2}; \ 3 \ 2 \ 1 \ +\frac{1}{2}; \ 3 \ 2 \ 1 \ -\frac{1}{2}; \ 3 \ 2 \ 0 \ +\frac{1}{2}; \ 3 \ 2 \ 0 \ -\frac{1}{2}; \ 3 \ 2 \ -2 \ +\frac{1}{2}; \ 3 \ 2 \ -2 \ -\frac{1}{2}; \ 3 \ 2 \ -1 \ +\frac{1}{2}; \ 3 \ 2 \ -1 \ -\frac{1}{2}$