7 The Quantum-Mechanical Model of the Atom

Review Questions

7.1 When a particle is absolutely small it means that you cannot observe it without disturbing it. When you observe the particle, it behaves differently than when you do not observe it. Electrons fit this description.

7.2 You can measure the position of a baseball by observing the light that strikes the ball, bounces off it, and enters your eye. The baseball is so large in comparison to the disturbance caused by the light that it is virtually unaffected by your observation. By contrast, if you attempt to measure the position of an electron using light, the light itself disturbs the electron. The interaction of the light with the electron changes its position.

7.3 The quantum-mechanical model of the atom is important because it explains how electrons exist in atoms and how those electrons determine the chemical and physical properties of elements.

7.4 Light is electromagnetic radiation, a type of energy embodied in oscillating electric and magnetic fields. Light in a vacuum travels at 3.00×10^8 m/s.

7.5 The wavelength (λ) of the wave is the distance in space between adjacent crests and is measured in units of distance. The amplitude of the wave is the vertical height of a crest. The more closely spaced the waves, that is, the shorter the wavelength, the more energy there is. The amplitude of the electric and magnetic field waves in light determine the intensity or brightness of the light. The higher the amplitude, the more energy the wave has.

7.6 The frequency, (ν), is the number of cycles (or wave crests) that pass through a stationary point in a given period of time. The units of frequency are cycles per second. The frequency is inversely proportional to the wavelength (λ). Frequency and wavelength are related by the equation $\nu = \frac{c}{\lambda}$.

7.7 For visible light, wavelength determines the color. Red light has a wavelength of 750 nm, the longest wavelength of visible light, and blue has a wavelength of 500 nm.

7.8 The presence of a variety of wavelengths in white light is responsible for the way we perceive colors in objects. When a substance absorbs some colors while reflecting others, it appears colored. Grass appears green because it reflects primarily the wavelength associated with green light and absorbs the others.

7.9 (a) Gamma rays(γ) – the wavelength range is 10^{-11} to 10^{-15} m. Gamma rays are produced by the sun, other stars, and certain unstable atomic nuclei on Earth. Human exposure to gamma rays is dangerous because the high energy of gamma rays can damage biological molecules.

(b) X-rays – the wavelength range is 10^{-8} to 10^{-11} m. X-rays are used in medicine. X-rays pass through many substances that block visible light and are therefore used to image bones and internal organs. X-rays are sufficiently energetic to damage biological molecules so, while several yearly exposures to X-rays are harmless, excessive exposure increases cancer risk.

threshold frequency below which no electrons were emitted from the metal, no matter how long the light shone on the metal. Low-frequency light would not eject electrons from a metal regardless of intensity or duration. But high-frequency light would eject electrons even at low intensity without any lag time.

7.13 Because of the results of the experiments with the photoelectric effect, Einstein proposed that light energy must come in packets. The amount of energy in a light packet depends on its frequency (wavelength). The emission of electrons depends on whether or not a single photon has sufficient energy to dislodge a single electron.

7.14 A photon is a packet of light. The energy of the photon can be expressed in terms of wavelength as $E = \frac{hc}{\lambda}$ or in terms of frequency as $E = h\nu$.

7.15 An emission spectrum occurs when an atom absorbs energy and re-emits that energy as light. The light emitted contains distinct wavelengths for each element. The emission spectrum of a particular element is always the same and can be used to identify the element. A white light spectrum is continuous, meaning that there are no sudden interruptions in the intensity of the light as a function of wavelengths. It consists of all wavelengths. Emission spectra are not continuous. They consist of bright lines at specific wavelengths, with complete darkness in between.

7.16 In the Bohr model electrons travel around the nucleus in circular orbits. Bohr's orbits could exist only at specific, fixed distances form the nucleus. The energy of each orbit was also fixed, or quantized. Bohr called these orbits stationary states and suggested that, although they obeyed the laws of classical mechanics, they also possessed "a peculiar, mechanically unexplainable, stability." Bohr further proposed that, in contradiction to classical electromagnetic theory, no radiation was emitted by an electron orbiting the nucleus in a stationary state. It was only when an electron jumped, or made a transition, from one stationary state to another that radiation was emitted or absorbed. The emission spectrum of an atom consisted of discrete lines because the stationary states existed only at specific, fixed energies. The energy of the photon created when an electron made a transition from one stationary state to another was simply the energy difference between the two stationary states.

7.17 Electron diffraction occurs when an electron beam is aimed at two closely spaced slits, and a series of detectors is arranged to detect the electrons after they pass through the slits. An interference pattern similar to that observed for light is recorded behind the slits. Electron diffraction is evidence of the wave nature of electrons.

7.18 The de Broglie wavelength is the wavelength associated with an electron traveling through space. It is related to its kinetic energy. The wavelength, λ, associated with an electron of mass, m, moving at velocity, v, is given by the de Broglie relation: $\lambda = \frac{h}{mv}$.

7.19 Complementary properties are those that exclude one another. The more you know about one, the less you know about the other. Which of two complementary properties you observe depends on the experiment you perform. In electron diffraction, when you try to observe which hole the electron goes through (particle nature) you lose the interference pattern (wave nature). When you try to observe the interference pattern, you cannot determine which hole the electron goes through.

7.20 Heisenberg's uncertainty principle states that the product of Δx and $m\Delta v$ must be greater than or equal to a finite number. In other words, the more accurately you know the position of an electron (the smaller Δx) the less accurately you can know its velocity (the bigger Δv) and vice versa. The complementarity of the wave nature and particle nature of the electron results in the complementarity of velocity and position. Heisenberg solved the contradiction of an object as both a particle and a wave by introducing complementarity—an electron is observed as either a particle or a wave, but never both at once.

7.21 A trajectory is a path that is determined by the particle's velocity (the speed and direction of travel), its position, and the forces acting on it. Both position and velocity are required to predict a trajectory.

7.22 Because the uncertainty principle says that you cannot know both the position and velocity of the electron simultaneously, you cannot predict the trajectory.

7.23 Deterministic means that the present determines the future. That means that under the identical condition, identical results will occur.

7.24 The indeterminate behavior of an electron means that under identical conditions, the electron does not have the same trajectory and does not "land" in the same spot each time.

7.25 A probability distribution map is a statistical map that shows where an electron is likely to be found under a given set of conditions.

7.26 Using the Schrödinger equation we describe the probability distribution maps for electron states. In these the electron has a well-defined energy, but not a well-defined position. In other words, for each state, we can specify the energy of the electron precisely, but not its location at a given instant. The electron's position is described in terms of an orbital.

7.27 An orbital is a probability distribution map showing where the electron is likely to be found.

7.28 The mathematical derivation of energies and orbitals for electrons in atoms comes from solving the Schrödinger equation. The general form of the Schrödinger equation is $\mathcal{H}\Psi = E\Psi$. The symbol $\mathcal{H}$ stands for the Hamiltonian operator, a set of mathematical operations that represent the total energy (kinetic and potential) of the electron within the atom. The symbol E is the actual energy of the electron. The symbol ψ is the wave function, a mathematical function that describes the wavelike nature of the electron. A plot of the wave function squared (Ψ^2) represents an orbital, a position probability distribution map of the electron.

7.29 The principal quantum number (n) is an integer and has possible values of 1,2,3, etc. The principal quantum number determines the overall size and energy of an orbital.

7.30 The angular momentum quantum number (l) is an integer and has possible values of 0,1,2,3, etc. The angular momentum quantum number determines the shape of the orbital.

7.31 The magnetic quantum number (m_l) is an integer ranging from $-l$ to $+l$. For example, if $l = 1$, $m_l = -1, 0, +1$. The magnetic quantum number specifies the orientation of the orbital.

7.32 (a) $n = 1, l = 0, m_l = 0$

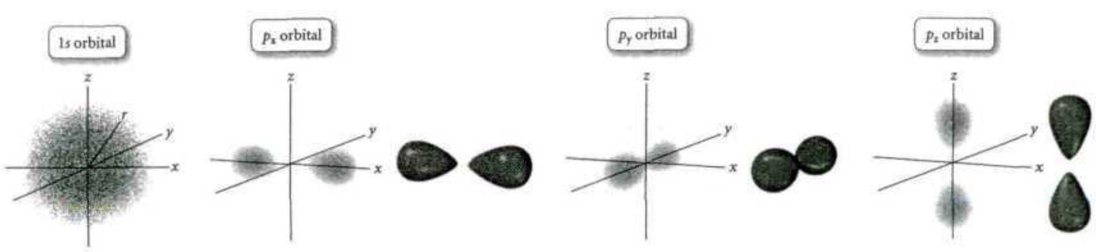
(b) $n = 2, l = 0, m_l = 0$; $n = 2, l = 1, m_l = -1$; $n = 2, l = 1, m_l = 0$; $n = 2, l = 1, m_l = +1$

(c) $n = 3, l = 0, m_l = 0$; $n = 3, l = 1, m_l = -1$; $n = 3, l = 1, m_l = 0$; $n = 3, l = 1, m_l = +1$; $n = 3, l = 2, m_l = -2$; $n = 3, l = 2, m_l = -1$; $n = 3, l = 2, m_l = 0$; $n = 3, l = 2, m_l = +1$; $n = 3, l = 2, m_l = +2$

(d) $n = 4, l = 0, m_l = 0$; $n = 4, l = 1, m_l = -1$; $n = 4, l = 1, m_l = 0$; $n = 4, l = 1, m_l = +1$; $n = 4, l = 2, m_l = -2$; $n = 4, l = 2, m_l = -1$; $n = 4, l = 2, m_l = 0$; $n = 4, l = 2, m_l = +1$; $n = 4, l = 2, m_l = +2$; $n = 4, l = 3, m_l = -3$; $n = 4, l = 3, m_l = -2$; $n = 4, l = 3, m_l = -1$; $n = 4, l = 3, m_l = 0$; $n = 4, l = 3, m_l = +1$; $n = 4, l = 3, m_l = +2$; $n = 4, l = 3, m_l = +3$

7.33 The probability density is the probability per unit volume of finding the electron at a point in space. The radial distribution function represents the total probability of finding the electron within a thin spherical shell at a distance r from the nucleus. In contrast to probability density, which has a maximum at the nucleus for an s orbital, the radial distribution function has a value of zero at the nucleus. It increases to a maximum and then decreases again with increasing r.

7.34



Check: The units of the answer, m, are correct. The magnitude of the answer is very small as would be expected for the de Broglie wavelength of a bullet.

7.57 **Given:** $\Delta x = 552\text{pm}$, $m = 9.109 \times 10^{-31}$ kg **Find:** Δv
Conceptual Plan: Δx, **m** $\rightarrow$ Δv

$$\Delta x \times m\Delta v \geq \frac{h}{4\pi}$$

Solution: $$\frac{6.626 \times 10^{-34} \frac{\cancel{kg} \cdot m^{\cancel{2}}}{s^{\cancel{2}}} \cdot \cancel{s}}{4(3.141)(9.109 \times 10^{-31}\ \cancel{kg})(552\ \cancel{pm})\left(\frac{\cancel{m}}{1 \times 10^{12}\ \cancel{pm}}\right)} = 1.05 \times 10^5\ \text{m/s}$$

Check: The units of the answer, m/s, are correct. The magnitude is reasonable for the uncertainty in the speed of an electron.

7.58 **Given:** $m = 9.109 \times 10^{-31}$ kg, $v = 3.7 \times 10^5$m/s, $\Delta v = 1.88 \times 10^5$m/s **Find:** Δx
Conceptual Plan Δv, **m** $\rightarrow$ Δx

$$\Delta x \times m\Delta v \geq \frac{h}{4\pi}$$

Solution: $$\frac{6.626 \times 10^{-34} \frac{\cancel{kg} \cdot \cancel{m}^{\cancel{2}}}{\cancel{s}^{\cancel{2}}} \cdot \cancel{s}\left(\frac{1 \times 10^{12}\ \text{pm}}{\cancel{m}}\right)}{4(3.141)(9.109 \times 10^{-31}\ \cancel{kg})\left(\frac{1.88 \times 10^5\ \cancel{m}}{\cancel{s}}\right)} = 308\ \text{pm}$$

Check: The units of the answer, pm, are correct. The magnitude is reasonable when compared to the speed of the electron.

Orbitals and Quantum Numbers

7.59 Since the size of the orbital is determined by the *n* quantum, with the size increasing with increasing *n*, an electron in a 2*s* orbital is closer, on average, to the nucleus than an electron in a 3*s* orbital.

7.60 Since the size of the orbital is determined by the *n* quantum, with the size increasing with increasing *n*, an electron in a 4*p* orbital is further away, on average, from the nucleus than an electron in a 3*p* orbital.

7.61 The value of *l* is an integer that lies between 0 and $n - 1$.

(a) When $n = 1$, *l* can only be $l = 0$.

(b) When $n = 2$, *l* can be $l = 0$ or $l = 1$.

(c) When $n = 3$, *l* can be $l = 0$, $l = 1$, or $l = 2$.

(d) When $n = 4$, *l* can be $l = 0$, $l = 1$, $l = 2$, or $l = 3$.

7.62 The value of m_l is an integer that lies between $-l$ and $+l$.

(a) When $l = 0$, m_l can only be $m_l = 0$.

(b) When $l = 1$, m_l can be $m_l = -1$, $m_l = 0$, or $m_l = +1$.

(c) When $l = 2$, m_l can be $m_l = -2$, $m_l = -1$, $m_l = 0$, $m_l = +1$, or $m_l = +2$.

(d) When $l = 3$, m_l can be $m_l = -3$, $m_l = -2$, $m_l = -1$, $m_l = 0$, $m_l = +1$, $m_l = +2$, or $m_l = +3$.

7.63 Set c cannot occur together as a set of quantum numbers to specify an orbital. *l* must lie between 0 and $n - 1$, so for $n = 3$, l can only be as high as 2.

7.64 (a) 1*s* is a real orbital, $n = 1$, $l = 0$.

(b) 2*p* is a real orbital, $n = 2$, $l = 1$.