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CHAPTER 12

QUANTUM MECHANICS AND ATOMIC THEORY

Discussion Questions

1. The equations relating the terms are $\nu\lambda = c$, $E = h\nu$, and $E = hc/\lambda$. From the equations, wavelength and frequency are inversely related, photon energy and frequency are directly related, and photon energy and wavelength are inversely related. In Figure 12.3, the EMR with the shortest wavelengths are gamma rays and the EMR with the longest wavelengths are AM radio waves. From the relationships discussed previously, gamma rays will have the highest frequency and the most energetic photons. Radio waves will have the lowest frequency and least energetic photons.

2. Wavelike properties refer to movement that is characterized by a wavelength and a frequency. Particulate properties refer to anything having a mass and a velocity along with a certain kinetic energy.

   Planck's discovery that heated bodies give off only certain frequencies of light and Einstein's study of the photoelectric effect support the quantum theory of light. The wave–particle duality is summed up by saying all matter exhibits both particulate and wave properties. Electromagnetic radiation, which was thought to be a pure waveform, transmits energy as if it has particulate properties. Conversely, electrons, which were thought to be particles, have a wavelength associated with them. This is true for all matter. Some evidence supporting wave properties of matter are:

   1. Electrons can be diffracted like light.
   2. The electron microscope uses electrons in a fashion similar to the way in which light is used in a light microscope.

   However, wave properties of matter are only important for small particles with a tiny mass (e.g., electrons). The wave properties of larger particles are not significant.

3. The H emission spectrum is just looking at the transitions in the visible region of the spectrum. If we were to look at the infrared or ultraviolet region, we would see a lot more transitions.

4. The Bohr model assumes that the electron in hydrogen can orbit the nucleus at specific distances from the nucleus. Each orbit has a specific energy associated with it. Therefore the electron in hydrogen can only have specific energies; not all energies are allowed. The term quantized refers to the allowed energy levels for the electron in hydrogen.

   The great success of the Bohr model is that it could explain the hydrogen emission spectrum. The electron in H moves about the allowed energy levels by absorbing or emitting certain photons of energy. The photon energies absorbed or emitted must exactly equal the energy

difference between any two allowed energy levels. Because not all energies are allowed in hydrogen (energy is quantized), not all energies of EMR are absorbed/emitted.

The Bohr model predicted the exact wavelengths of light that would be emitted for a hydro-gen atom. Although the Bohr model has great success for hydrogen and other one-electron ions, it does not explain emission spectra for elements/ions having more than one electron. A fundamental flaw is that we cannot know the exact motion of an electron as it moves about the nucleus, therefore well-defined circular orbits are not appropriate. Another flaw is that it does not consider the electron–electron repulsions in atoms/ions containing more than one electron.

5. Mg: $1s^2 2s^2 2p^6 3s^2$; the first ionization energy (IE) for Mg refers to losing one of the 3s electrons. The effective nuclear charge is the apparent nuclear charge exerted on a particular electron. Assuming the 10 inner core electrons in the 1s, 2s, and 2p orbitals completely shield the 3s valence electrons from the actual 12+ charge from the nucleus, then $Z_{eff} = 12 - 10 = +2$.

To calculate Z_{eff}, let's assume the Bohr model applies to the 3s electron, $E_{3s} = -R_H Z^2/n^2 = -R_H Z^2_{eff}/n_2$, where $n = 3$. $IE = E_\infty - E_{3s} = 0 - E_{3s} = R_H Z^2_{eff}/3^2$.

$$\frac{735\ \text{kJ}}{\text{mol}} \times \frac{1\ \text{mol}}{6.022 \times 10^{23}\ \text{Mg atoms}} \times \frac{1000\ \text{J}}{\text{kJ}} = \frac{2.178 \times 10^{-18}\ \text{J}\ \left(Z_{eff}\right)^2}{3^2},\ Z_{eff} = 2.25$$

The calculated Z_{eff} is larger than the estimated Z_{eff} from the simplistic assumption that inner core electrons completely shield the valence electrons from the nuclear charge; this assumption is not valid. At any one time, any inner core electron has a small probability of being further from the nucleus than the valence electron. In this case, the shielding electrons are something less than the number of inner core electrons. For Mg, the number of shielding electrons is not 10, but something smaller than 10. This gives a Z_{eff} that is larger than +2.

6. Each element shows a very large energy change between the second and third IEs. This suggests that both elements have two valence electrons, so they are alkaline earth metals. Because all ionization IEs for element Y are larger than those of element X, element Y should be above element X in the periodic table. One possible example is Y = Mg and X = Ca. Of course, there are other combinations from Group 2A that would work.

7. Consecutive IEs always increase. So, the second IE of He would be greater than the first IE. For the first IE, helium has two electrons attracted to a nucleus containing two protons. For the second IE, there is just a single electron attracted to the two protons in the nucleus. The nucleus exerts a stronger attraction when there are fewer electrons present. Hence the second IE is larger than the first IE.

8. Li: $1s^2 2s^1$; Be: $1s^2 2s^2$: The second IE for beryllium removes an electron from the valence n = 2 level. The second IE for Li removes an electron from an inner core n = 1 level. Inner core electrons are much closer to the nucleus, on average, and are significantly more attracted to the nucleus than valence electrons. Hence the second IE for lithium where an inner core electron is removed should be larger.

9. As atomic number increases across a row, IE generally increases. There are two major exceptions to this trend for elements 1–36. There is a decrease in ionization when going from a 2A to a 3A element and from a 5A to a 6A element. To explain these, let's look at specific examples in row 3. Aluminum is out of order because the electrons in the filled 3s orbital shield some of the nuclear charge from the 3p electron. Hence the 3p electron is less tightly bound than a 3s electron, resulting in a lower IE for aluminum as compared to magnesium. The IE of sulfur is lower than phosphorus because of the extra electron–electron repulsions in the doubly occupied sulfur 3p orbital. These added repulsions, which are not present in phosphorus, make it slightly easier to remove an electron from sulfur as compared to phosphorus.

10. The second IE for sodium removes an inner core electron. The second IE for the remaining elements in row 3 remove valence electrons. Inner core electrons always require significantly more energy to remove as compared to valence electrons. So sodium will have the largest second IE. We would expect the general trend to apply for the remaining elements in row 3. That is, we would expect there to be an increase in the second IE from magnesium to argon (with perhaps two exceptions to mirror the first IE trend exceptions).

11. There are several possibilities for a rational answer. One possibility is that nonmetals generally have smaller size and larger IEs. As one goes vertically down the periodic table, size increases while IE increases. This results in a decrease in nonmetallic properties as one goes down a group. As one goes across the periodic table from left to right, size decreases and IE increases. This results in nonmetallic behavior increasing left to right across a periodic table. Combining these two trends, one would expect the line dividing metals from nonmetals to be a diagonal line going downward from left to right. This is what is required to have the correct nonmetal combination of relatively small size and relatively large IEs.

12. An electron has wave and particle properties, and its energy is quantized. An electron's exact location in an atom can never be known. Instead, we talk about probabilities of where electrons are in the various allowed energy levels (orbitals). Since we cannot know specifically where the electrons are in an atom, measurements of size (atomic radius) cannot be determined exactly. In terms of the quantum mechanical model, the concept of an electron is vague, not the typical negative-charged particle with a known orbit concept.

13. $Cl(g) \rightarrow Cl^+(g) + e^-$ ΔH = IE for Cl; the electron affinity (EA) energy changes for choices a–d are:

a. $Cl(g) + e^- \rightarrow Cl^-(g)$; b. $Cl^- + e^- \rightarrow Cl^{2-}$; c. $Cl^+ + e^- \rightarrow Cl$; d. $F + e^- \rightarrow F^-$

The EA of Cl^+ (answer c) is the exact opposite of the IE for Cl. So, these two processes are equal in magnitude but opposite in sign. The IE for Cl will be positive while the EA for Cl^+ will be negative.

14. Electrons in an atom are attracted to the nucleus of that atom. To remove an electron, energy must be added to break this attraction. So, IEs are never negative quantities. It is correct that

when K loses an electron it does obtain a noble gas electron configuration. But it does not lose an electron unless energy is added.

15. The effective nuclear charge, Z_{eff}, is the apparent nuclear charge exerted on a particular electron. The nuclear charge, Z, is the number of protons in the nucleus. For any electron from a polyelectronic substance, there will be electron–electron repulsions. These repulsions reduce the nuclear charge experienced by a particular electron. Substances that only have one electron like H, He^+, Li^{2+}, etc., do not have electron repulsions. For one electron species, $Z_{eff} = Z$.

16. In going down a column, the added electrons go into higher n value orbitals. These higher n value orbitals have the electrons further from the nucleus, on average, resulting in a larger atomic size and smaller IEs. As electrons are added going across the periodic table, the added electrons go into the same n value orbitals. However, electrons in the same n value orbitals don't completely shield the increasing nuclear charge caused by the added protons. This results in an increase in effective nuclear charge as one goes across the periodic table, so the electrons are held more strongly and IE increases (size decreases).

17. A probability density distribution is the square of the wave function for an orbital indicating the probability of finding an electron in that orbital at a particular point in space. This probability is greatest close to the nucleus and drops off rapidly as the distance from the nucleus increases.

The radial probability distribution represents the total probability of finding the electron at a particular distance from the nucleus in an orbital. For the distribution in the 1s wave function, the 1s orbital is divided into successive thin spherical shells and the total probability of finding the electron in each spherical shell is plotted versus distance from the nucleus.

18. With the higher nuclear charge of Li, one would expect the Li 1s orbital to be lower in energy than the H 1s orbital. Any electron placed in a Li 1s orbital will be attracted to a larger positive charge from the nucleus and will be more stable, which corresponds to a lower potential energy. In general, the zero-energy state is when an electron is infinitely far away from the nucleus so that it exhibits no attraction. When the electron in an orbital is close enough to be attracted to the nucleus, the potential energy is reduced by this attraction. With the defined zero state representing no attraction between the electrons and the nucleus, an orbital that has electrons that are attracted to the nucleus will have a negative energy value. The lowest energy state has the electron in that state with the greatest attraction to the nucleus; this is also the most stable state.

19. One would expect the Li 1s orbital to be smaller than the H 1s orbital because of the greater nuclear charge for Li. In terms of size, H is smaller than Li. Li has a 2s electron in addition to the 1s electrons. The 2s electron for Li will be on average further from the nucleus than the 1s electron for H, making Li larger than H.

20. Hydrogen has all the various orbitals, including a 3s orbital. For hydrogen, the energy of a particular orbital is just determined by *n*. Thus a 3s orbital is degenerate (has the same energy) as the 3p orbitals, which have the same energy as the 3d orbitals.

21. In hydrogen, a 2s and 2p orbital have the same energy. For hydrogen, the energy of a particular orbital is just determined by *n*. In species with more than one electron (like helium), the 2s orbital has a lower energy than the 2p orbitals. In polyelectronic species, the *n*s orbitals have lower energy than the *n*p orbitals, which are lower than the *n*d orbitals, etc.

22. The IE for sodium is 495 kJ/mol and the IE for fluorine is 1681 kJ. Each of these values represents the energy it takes for any substance to remove an electron from that atom (in the gaseous state). Since sodium has the smaller IE, then it is easier to remove an electron from sodium than from fluorine.

23. The first period corresponds to $n = 1$, which can only have 1s orbitals. The 1s orbital could hold 3 electrons, hence the first period would have three elements. The second period corresponds to $n = 2$, which has 2s and 2p orbitals. These four orbitals can each hold three electrons. A total of 12 elements would be in the second period. The third period has 3s and 3p orbitals, so the third row will also have 12 elements. The first three rows of the periodic table will have 3 + 12 +12 = 27 elements instead of 18 as we have in our current periodic table. A sketch of the s and p block elements for the first three rows follows.

1	2																3
4	5	6	7	8	9	10	11	12	13	14	15						
16	17	18	19	20	21	22	23	24	25	26	27						

The first three noble gases would be 3, 15, and 27. The next noble gas would have 4s, 3d, and 4p orbitals. If each of the five 3d orbitals could hold 3 electrons, the first row of the transition metals would be 15 elements long. This makes the fourth row of the periodic table 27 elements long versus the current 18 elements long. The next noble gas after 27 will be 54, and the next noble gas after 54 will be 51 + 27 = 81. The last orbitals filled for the actinide series are 5f orbitals. With 7 degenerate 5f orbitals and 3 electrons in each orbital, the actinide series will be 21 elements long.

24. The third IE refers to the following process: $E^{2+}(g) \rightarrow E^{3+}(g) + e^- \quad \Delta H = I_3$. The electron configurations for the 2+ charged ions of Na to Ar are:

Na^{2+}: $1s^2 2s^2 2p^5$

Mg^{2+}: $1s^2 2s^2 2p^6$

Al^{2+}: $[Ne]3s^1$

Si^{2+}: $[Ne]3s^2$

P^{2+}: $[Ne]3s^2 3p^1$

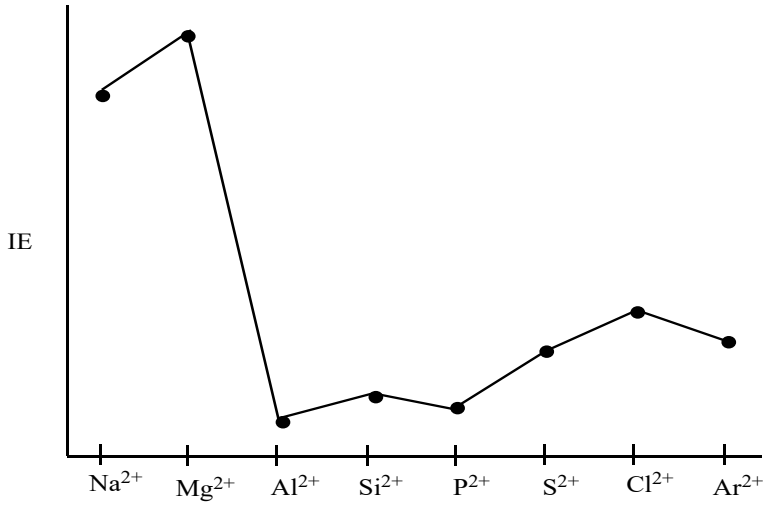
S^{2+}: $[Ne]3s^2 3p^2$

Cl^{2+}: $[Ne]3s^2 3p^3$

Ar^{2+}: $[Ne]3s^2 3p^4$

I_3 for sodium and magnesium should be extremely large compared with the others because n = 2 electrons are much more difficult to remove than n = 3 electrons. Between Na^{2+} and Mg^{2+}, one would expect to have the same trend as seen with I_1(F) versus I_1(Ne); these neutral atoms have identical electron configurations to Na^{2+} and Mg^{2+}. Therefore the $1s^2 2s^2 2p^5$ ion (Na^{2+}) should have a lower IE than the $1s^2 2s^2 2p^6$ ion (Mg^{2+}).

The remaining 2+ ions (Al^{2+} to Ar^{2+}) should follow the same trend as the neutral atoms having the same electron configurations. The general IE trend predicts an increase from $[Ne]3s^1$ to $[Ne]3s^2 3p^4$. The exceptions occur between $[Ne]3s^2$ and $[Ne]3s^2 3p^1$ and between $[Ne]3s^2 3p^3$ and $[Ne]3s^2 3p^4$. $[Ne]3s^2 3p^1$ is out of order because of the small penetrating ability of the 3p electron as compared with the 3s electrons. $[Ne]3s^2 3p^4$ is out of order because of the extra electron–electron repulsions present when two electrons are paired in the same orbital. Therefore the correct ordering for Al^{2+} to Ar^{2+} should be $Al^{2+} < P^{2+} < Si^{2+} < S^{2+} < Ar^{2+} < Cl^{2+}$, where P^{2+} and Ar^{2+} are out of line for the same reasons that Al and S are out of line in the general IE trend for neutral atoms. A qualitative plot of the third IEs for elements Na through Ar follows.



Note: The actual numbers in Table 12.6 support most of this plot. No I_3 is given for Na^{2+}, so you cannot check this. The only deviation from our discussion is I_3 for Ar^{2+}, which is greater than I_3 for Cl^{2+} instead of less than.

Light and Matter

25. The equations relating the terms are $\nu\lambda = c$, $E = h\nu$, and $E = hc/\lambda$. Frequency is defined as the number of wave cycles that pass a certain part per second, so statement a is true. Statement c is true because wavelength and frequency are inversely related. Statements b and d are false. All

electromagnetic radiation travels at the same speed, the speed of light, c. And frequency and photon energy are directly related. As the frequency is halved, photon energy is halved.

26. $\lambda = \dfrac{c}{\nu} = \dfrac{3.00 \times 10^{8}\ \text{m/s}}{4.5 \times 10^{14}\ \text{s}^{-1}} = 6.7 \times 10^{-7}\ \text{m}$; from Figure 12.3, this is red visible light.

27. $\nu = \dfrac{c}{\lambda} = \dfrac{3.00 \times 10^{8}\ \text{m/s}}{1.0 \times 10^{-2}\ \text{m}} = 3.0 \times 10^{10}\ \text{s}^{-1}$

$E = h\nu = 6.63 \times 10^{-34}\ \text{J s} \times 3.0 \times 10^{10}\ \text{s}^{-1} = 2.0 \times 10^{-23}\ \text{J/photon}$

$$\frac{2.0 \times 10^{-23}\ \text{J}}{\text{photon}} = \frac{6.02 \times 10^{23}\ \text{photons}}{\text{mol}} = 12\ \text{J/mol}$$

28. The wavelength is the distance between consecutive wave peaks. Wave *a* shows 4 wave-lengths and wave *b* shows 8 wavelengths.

Wave *a*: $\lambda = \dfrac{1.6 \times 10^{-3}\ \text{m}}{4} = 4.0 \times 10^{-4}\ \text{m}$

Wave *b*: $\lambda = \dfrac{1.6 \times 10^{-3}\ \text{m}}{8} = 2.0 \times 10^{-4}\ \text{m}$

Wave *a* has the longer wavelength. Because frequency and photon energy are both inversely proportional to wavelength, wave *b* will have the higher frequency and larger photon energy since it has the shorter wavelength.

$$\nu = \frac{c}{\lambda} = \frac{3.00 \times 10^{8}\ \text{m/s}}{2.0 \times 10^{-4}\ \text{m}} = 1.5 \times 10^{12}\ \text{s}^{-1}$$

$$E = \frac{hc}{\lambda} = \frac{6.63 \times 10^{-34}\ \text{J s} \times 3.00 \times 10^{8}\ \text{m/s}}{2.0 \times 10^{-4}\ \text{m}} = 9.9 \times 10^{-22}\ \text{J}$$

Because both waves are examples of electromagnetic radiation, both waves travel at the same velocity, c, the speed of light. From Figure 12.3 of the text, both of these waves represent infrared electromagnetic radiation.

29. 280 nm: $\nu = \dfrac{c}{\lambda} = \dfrac{3.00 \times 10^{8}\ \text{m/s}}{280\ \text{nm} \times \dfrac{1\ \text{m}}{1 \times 10^{9}\ \text{nm}}} = 1.1 \times 10^{15}\ \text{s}^{-1}$

320 nm: $\nu = \dfrac{3.00 \times 10^{8}\ \text{m/s}}{320 \times 10^{-9}\ \text{nm}} = 9.4 \times 10^{14}\ \text{s}^{-1}$

The compounds in the sunscreen absorb ultraviolet B (UVB) electromagnetic radiation having a frequency from $9.4 \times 10^{14}\ s^{-1}$ to $1.1 \times 10^{15}\ s^{-1}$.

30. S-type cone receptors: $\lambda = \dfrac{c}{\nu} = \dfrac{2.998 \times 10^8\ m/s}{6.00 \times 10^{14}\ s^{-1}} = 5.00 \times 10^{-7}\ m = 500.\ nm$

$$\lambda = \frac{2.998 \times 10^8\ m/s}{7.49 \times 10^{14}\ s^{-1}} = 4.00 \times 10^{-7}\ m = 400.\ nm$$

S-type cone receptors detect 400–500 nm light. From Figure 12.3 in the text, this is violet to green light, respectively.

M-type cone receptors: $\lambda = \dfrac{2.998 \times 10^8\ m/s}{4.76 \times 10^{14}\ s^{-1}} = 6.30 \times 10^{-7}\ m = 630.\ nm$

$$\lambda = \frac{2.998 \times 10^8\ m/s}{6.62 \times 10^{14}\ s^{-1}} = 4.53 \times 10^{-7}\ m = 453\ nm$$

M-type cone receptors detect 450–630 nm light. From Figure 12.3 in the text, this is blue to orange light, respectively.

L-type cone receptors: $\lambda = \dfrac{2.998 \times 10^8\ m/s}{4.28 \times 10^{14}\ s^{-1}} = 7.00 \times 10^{-7}\ m = 700.\ nm$

$$\lambda = \frac{2.998 \times 10^8\ m/s}{6.00 \times 10^{14}\ s^{-1}} = 5.00 \times 10^{-7}\ m = 500.\ nm$$

L-type cone receptors detect 500–700 nm light; this represents green to red light, respectively.

31. Referencing Figure 12.3 of the text, 2.12×10^{-10} m electromagnetic radiation is X rays.

$$\lambda = \frac{c}{\nu} = \frac{2.9979 \times 10^8\ m/s}{107.1 \times 10^6\ s^{-1}} = 2.799\ m$$

From the wavelength calculated above, 107.1 MHz electromagnetic radiation is FM radio waves.

$$\lambda = \frac{hc}{E} = \frac{6.626 \times 10^{-34}\ J\,s \times 2.998 \times 10^8\ m/s}{3.97 \times 10^{-19}\ J} = 5.00 \times 10^{-7}\ m$$

The 3.97×10^{-19} J/photon electromagnetic radiation is visible (green) light.

The photon energy and frequency order will be the exact opposite of the wavelength ordering because E and ν are both inversely related to λ. From the previously calculated wavelengths, the order of photon energy and frequency is:

FM radio waves < visible (green) light < X rays

longest λ	shortest λ
lowest ν	highest ν
smallest E	largest E

32. $E = \dfrac{310.\text{ kJ}}{\text{mol}} \times \dfrac{1\text{ mol}}{6.022 \times 10^{23}} = 5.15 \times 10^{-22}\text{ kJ} = 5.15 \times 10^{-19}\text{ J}$

$E = \dfrac{hc}{\lambda}, \quad \lambda = \dfrac{hc}{E} = \dfrac{6.626 \times 10^{-34}\text{ J s} \times 2.998 \times 10^{8}\text{ m/s}}{5.15 \times 10^{-19}\text{ J}} = 3.86 \times 10^{-7}\text{ m} = 386\text{ nm}$

33. a. $\lambda = \dfrac{c}{\nu} = \dfrac{3.00 \times 10^{8}\text{ m/s}}{6.0 \times 10^{13}\text{ s}^{-1}} = 5.0 \times 10^{-6}\text{ m}$

b. From Figure 12.3, this is infrared EMR.

c. $E = h\nu = 6.63 \times 10^{-34}\text{ J s} \times 6.0 \times 10^{13}\text{ s}^{-1} = 4.0 \times 10^{-20}\text{ J/photon}$

$\dfrac{4.0 \times 10^{-20}\text{ J}}{\text{photon}} \times \dfrac{6.022 \times 10^{23}\text{ photons}}{\text{mol}} = 2.4 \times 10^{4}\text{ J/mol}$

d. Frequency and photon energy are directly related ($E = h\nu$). Because $5.4 \times 10^{13}\text{ s}^{-1}$ EMR has a lower frequency than $6.0 \times 10^{13}\text{ s}^{-1}$ EMR, the $5.4 \times 10^{13}\text{ s}^{-1}$ EMR will have less energetic photons.

34. $E_{photon} = h\nu = \dfrac{hc}{\lambda}, \quad E_{photon} = \dfrac{6.626 \times 10^{-34}\text{ J s} \times 2.998 \times 10^{8}\text{ m/s}}{1.0 \times 10^{-10}\text{ m}} = 2.0 \times 10^{-15}\text{ J}$

$\dfrac{2.0 \times 10^{-15}\text{ J}}{\text{photon}} \times \dfrac{6.02 \times 10^{23}\text{ photons}}{\text{mol}} \times \dfrac{1\text{ kJ}}{1000\text{ J}} = 1.2 \times 10^{6}\text{ kJ/mol}$

$E_{photon} = \dfrac{6.626 \times 10^{-34}\text{ J s} \times 2.998 \times 10^{8}\text{ m/s}}{1.0 \times 10^{4}\text{ m}} = 2.0 \times 10^{-29}\text{ J}$

$\dfrac{2.0 \times 10^{-29}\text{ J}}{\text{photon}} \times \dfrac{6.02 \times 10^{23}\text{ photons}}{\text{mol}} \times \dfrac{1\text{ kJ}}{1000\text{ J}} = 1.2 \times 10^{-8}\text{ kJ/mol}$

X rays do have energy greater than the carbon–carbon bond energy, therefore X rays could conceivably break carbon–carbon bonds in organic compounds and thereby disrupt the function of an organic molecule. Radio waves, however, do not have sufficient energy to break carbon–carbon bonds and are therefore relatively harmless.

35. The energy needed to remove a single electron is:

$$\frac{279.7\ \text{kJ}}{\text{mol}} \times \frac{1\ \text{mol}}{6.0221 \times 10^{23}} = 4.645 \times 10^{-22}\ \text{kJ} = 4.645 \times 10^{-19}\ \text{J}$$

$$E = \frac{hc}{\lambda},\ \lambda = \frac{hc}{E} = \frac{6.6261 \times 10^{-34}\ \text{J s} \times 2.9979 \times 10^{8}\ \text{m/s}}{4.645 \times 10^{-19}\ \text{J}} = 4.277 \times 10^{-7}\ \text{m} = 427.7\ \text{nm}$$

36. $$\frac{890.1\ \text{kJ}}{\text{mol}} \times \frac{1\ \text{mol}}{6.0221 \times 10^{23}\ \text{atoms}} = \frac{1.478 \times 10^{-21}\ \text{kJ}}{\text{atom}} = \frac{1.478 \times 10^{-18}\ \text{J}}{\text{atom}}$$

= IE per atom

$$E = \frac{hc}{\lambda},\ \lambda = \frac{hc}{E} = \frac{6.626 \times 10^{-34}\ \text{J s} \times 2.9979 \times 10^{8}\ \text{m/s}}{1.478 \times 10^{-18}\ \text{J}} = 1.344 \times 10^{-7}\ \text{m} = 134.4\ \text{nm}$$

No, it will take light having a wavelength of 134.4 nm or less to ionize gold. A photon of light having a wavelength of 225 nm is longer wavelength and thus lower energy than 134.4 nm light.

37. The energy to remove a single electron is:

$$\frac{208.4\ \text{kJ}}{\text{mol}} \times \frac{1\ \text{mol}}{6.022 \times 10^{23}} = 3.461 \times 10^{-22}\ \text{kJ} = 3.461 \times 10^{-19}\ \text{J} = E_w$$

Energy of 254-nm light is:

$$E = \frac{hc}{\lambda} = \frac{(6.626 \times 10^{-34}\ \text{J s})(2.998 \times 10^{8}\ \text{m/s})}{254 \times 10^{-9}\ \text{m}} = 7.82 \times 10^{-19}\ \text{J}$$

$E_{photon} = E_K + E_w$, $E_K = 7.82 \times 10^{-19}\ \text{J} - 3.461 \times 10^{-19}\ \text{J} = 4.36 \times 10^{-19}\ \text{J}$ = maximum KE

38. Planck's discovery that heated bodies give off only certain frequencies of light and Einstein's study of the photoelectric effect support the quantum theory of light. The wave–particle duality is summed up by saying all matter exhibits both particulate and wave properties. Electromagnetic radiation, which was thought to be a pure waveform, transmits energy as if it has particulate properties. Conversely, electrons, which were thought to be particles, have a wavelength associated with them. This is true for all matter. Some evidence supporting wave properties of matter are:

1. Electrons can be diffracted like light.
2. The electron microscope uses electrons in a fashion similar to the way in which light is used in a light microscope.

However, wave properties of matter are only important for small particles with a tiny mass (e.g., electrons). The wave properties of larger particles are not significant.

39. The photoelectric effect refers to the phenomenon in which electrons are emitted from the surface of a metal when light strikes it. The light must have a certain minimum frequency (energy) to

remove electrons from the surface of a metal. Light having a frequency below the minimum results in no electrons being emitted, whereas light at or higher than the minimum frequency does cause electrons to be emitted. For light having a frequency higher than the minimum frequency, the excess energy is transferred into kinetic energy for the emitted electron. Albert Einstein explained the photoelectric effect by applying quantum theory.

40. a. 10.% of speed of light = $0.10 \times 3.00 \times 10^8$ m/s = 3.0×10^7 m/s

$$\lambda = \frac{h}{mv},\ \lambda = \frac{6.63 \times 10^{-34}\ \text{J s}}{9.11 \times 10^{-31}\ \text{kg} \times 3.0 \times 10^7\ \text{m/s}} = 2.4 \times 10^{-11}\ \text{m} = 2.4 \times 10^{-2}\ \text{nm}$$

Note: For units to come out, the mass must be in kg since 1 J = 1 kg m^2/s^2.

b. $$\lambda = \frac{h}{mv} = \frac{6.63 \times 10^{-34}\ \text{J s}}{0.055\ \text{kg} \times 35\ \text{m/s}} = 3.4 \times 10^{-34}\ \text{m} = 3.4 \times 10^{-25}\ \text{nm}$$

This number is so small that it is insignificant. We cannot detect a wavelength this small. The meaning of this number is that we do not have to worry about the wave properties of large objects.

41. a. $$\lambda = \frac{h}{mv} = \frac{6.626 \times 10^{-34}\ \text{J s}}{1.675 \times 10^{-27}\ \text{kg} \times (0.0100 \times 2.998 \times 10^8\ \text{m/s})} = 1.32 \times 10^{-13}\ \text{m}$$

b. $$\lambda = \frac{h}{mv},\ v = \frac{h}{\lambda m} = \frac{6.626 \times 10^{-34}\ \text{J s}}{75 \times 10^{-12}\ \text{m} \times 1.675 \times 10^{-27}\ \text{kg}} = 5.3 \times 10^3\ \text{m/s}$$

42. $\lambda = \frac{h}{mv},\ v = \frac{h}{m\lambda}$; for $\lambda = 0.1\ \text{nm} = 1 \times 10^{-10}$ m:

$$v = \frac{6.626 \times 10^{-34}\ \text{J s}}{9.11 \times 10^{-31}\ \text{kg} \times 1 \times 10^{-10}\ \text{m}} = 7 \times 10^6\ \text{m/s}$$

43. $$m = \frac{h}{\lambda v} = \frac{6.626 \times 10^{-34}\ \text{kg m}^2/\text{s}}{3.31 \times 10^{-15}\ \text{m} \times (0.0100 \times 2.998 \times 10^8\ \text{m/s})} = 6.68 \times 10^{-26}\ \text{kg/atom}$$

$$\frac{6.68 \times 10^{-26}\ \text{kg}}{\text{atom}} \times \frac{6.022 \times 10^{23}\ \text{atoms}}{\text{mol}} \times \frac{1000\ \text{g}}{\text{kg}} = 40.2\ \text{g/mol}$$

The element is calcium (Ca).

Hydrogen Atom: The Bohr Model

44. The Bohr model assumes that the electron in hydrogen can orbit the nucleus at specific distances from the nucleus. Each orbit has a specific energy associated with it. Therefore the electron in hydrogen can only have specific energies; not all energies are allowed. The term quantized refers to the allowed energy levels for the electron in hydrogen.

The great success of the Bohr model is that it could explain the hydrogen emission spectrum. The electron in H moves about the allowed energy levels by absorbing or emitting certain photons of energy. The photon energies absorbed or emitted must be exactly equal to the energy difference between any two allowed energy levels. Because not all energies are allowed in hydrogen (energy is quantized), not all energies of EMR are absorbed/emitted.

The Bohr model predicted the exact wavelengths of light that would be emitted for a hydrogen atom. Although the Bohr model has great success for hydrogen and other one-electron ions, it does not explain emission spectra for elements/ions having more than one electron. The fundamental flaw is that we cannot know the exact motion of an electron as it moves about the nucleus, therefore well-defined circular orbits are not appropriate.

45. a. For hydrogen (Z = 1), the energy levels in units of joules are given by the equation $E_n = -2.178 \times 10^{-18}(1/n^2)$. As n increases, the differences between $1/n^2$ for consecutive energy levels becomes smaller and smaller. Consider the difference between $1/n^2$ values for $n = 1$ and $n = 2$ as compared to $n = 3$ and $n = 4$.

For $n = 1$ and $n = 2$: For $n = 3$ and $n = 4$:

$$\frac{1}{1^2} - \frac{1}{2^2} = 1 - 0.25 = 0.75 \qquad \frac{1}{3^2} - \frac{1}{4^2} = 0.1111 - 0.0625 = 0.0486$$

Because the differences between $1/n^2$ values for consecutive energy levels decrease as n increases, the energy levels get closer together as n increases.

b. For a spectral transition for hydrogen, $\Delta E = E_f - E_i$:

$$\Delta E = -2.178 \times 10^{-18}\ \text{J}\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$

where n_i and n_f are the levels of the initial and final states, respectively. A positive value of ΔE always corresponds to an absorption of light, and a negative value of ΔE always corresponds to an emission of light.

In the diagram, the red line is for the $n_i = 3$ to $n_f = 2$ transition.

$$\Delta E = -2.178 \times 10^{-18}\ \text{J}\left(\frac{1}{2^2} - \frac{1}{3^2}\right) = -2.178 \times 10^{-18}\ \text{J}\left(\frac{1}{4} - \frac{1}{9}\right)$$

$$\Delta E = -2.178 \times 10^{-18}\ \text{J} \times (0.2500 - 0.1111) = -3.025 \times 10^{-19}\ \text{J}$$

The photon of light must have precisely this energy (3.025×10^{-19} J).