Chapter 2

Quantum Theory

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1. **Which has more energy, a photon of microwave radiation or a photon of infrared radiation?**

Infrared radiation has a shorter wavelength and a higher frequency than microwave radiation (Figure 2.2). Consequently, infrared photons are more energetic than microwave photons.

3. **Radiation from the sun can cause the decomposition of O_2 in the stratosphere: $O_2 \rightarrow 2O$. This requires 495 kJ/mol of energy. What is the longest wavelength of light in nm that can accomplish this process? In what region of the electromagnetic spectrum (*i.e.*, infrared, visible, ultraviolet, etc.) is light of this energy found?**

$$E = \frac{4.95 \times 10^5 \text{ J}}{\text{mol}} \times \frac{1 \text{ mol}}{6.02 \times 10^{23} \text{ photon}} = 7.72 \times 10^{-19} \text{ J/photon}$$

We obtain the frequency of a photon with this energy from Equation 2.1

$$\nu = \frac{E}{h} = \frac{7.72 \times 10^{-19} \text{J/photon}}{6.63 \times 10^{-34} \text{ J} \times \text{s}} = 1.17 \times 10^{15} \text{ s}^{-1}$$

The wavelength of light with this frequency is

$$\lambda = \frac{c}{\nu} = \frac{3.00 \times 10^8 \text{ m} \cdot \text{s}^{-1}}{1.17 \times 10^{15} \text{ s}^{-1}} = 2.56 \times 10^{-7} \text{ m} = 256 \text{ nm}$$

a UV (ultraviolet) photon

5. **Fill in the following table:**

λ (nm)	ν (s^{-1})	E/photon (J)	E/mol photons (kJ/mol)
333	9.00×10^{14}	5.97×10^{-19}	360
47	$\mathbf{6.42 \times 10^{15}}$	4.25×10^{-18}	2.56×10^3
107	2.81×10^{15}	$\mathbf{1.86 \times 10^{-18}}$	1.12×10^3
1.11×10^3	2.71×10^{14}	1.79×10^{-19}	**108**

7. **Use the Bohr model to determine the radii of the n = 2 and n = 4 orbits of a He^{1+} ion.**

$$r_2 = 5.292 \times 10^{-11} \left(\frac{2^2}{2} \right) = 1.058 \times 10^{-11} \text{ meters} = 10.58 \text{ pm}$$

Use Equation 2.4 and Z = 2:

$$r_4 = 5.292 \times 10^{-11} \left(\frac{4^2}{2} \right) = 4.236 \times 10^{-10} \text{ meters} = 423.6 \text{ pm}$$

9. **Determine the energies of the n = 2 and n = 3 levels of a He^{1+} ion. What region of the spectrum (*i.e.*, infrared, visible, ultraviolet, etc.) would the n = 3 $\rightarrow$ 2 transition in He^{1+} ion occur? Is this transition an absorption or an emission?**

Use Z = 2 in Equation 2.5: $-E_n = -2.180 \times 10^{-18} \left(\frac{Z^2}{n^2} \right)$

$$E_2 = -2.180 \times 10^{-18} \left(\frac{2^2}{2^2} \right) = -2.180 \times 10^{-18} \text{ J} \qquad E_3 = -2.180 \times 10^{-18} \left(\frac{2^2}{3^2} \right) = -9.689 \times 10^{-19} \text{ J}$$

$$\Delta E = E_3 - E_2 = -9.689 \times 10^{-19} \text{ J} + 2.180 \times 10^{-18} \text{ J} = 1.21 \times 10^{-18} \text{ J} = h\nu$$

$$\nu = \frac{\Delta E}{h} = \frac{1.21 \times 10^{-18} \text{ J}}{6.626 \times 10^{-34} \text{ J} \cdot \text{s}} = 1.83 \times 10^{15} \text{ s}^{-1}$$

a photon of this frequency is in the ultraviolet (violet light ~ $7.5 \times 10^{14} \text{ s}^{-1}$)

11. **A 20 cm string is fastened at both sides is plucked. What is the wavelength of the n = 5 standing wave? How many nodes does it contain? Draw the wave.**

There are n-1 = 4 nodes, which must be situated symmetrically. Thus, the nodes are at 4, 8, 12, and 16 cm. $\lambda = 2L/n = 2(20)/5 = 8$ cm

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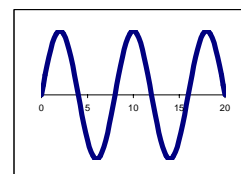
Quantum Theory

13. Which restriction on the quantum numbers is violated in the following sets:

n	l	m_l	m_s	
2	1	-1	<u>1</u>	$m_s \neq \pm \frac{1}{2}$
3	<u>-2</u>	0	$\frac{1}{2}$	$l < 0$
2	<u>2</u>	0	$-\frac{1}{2}$	$l = n$
3	2	<u>3</u>	$-\frac{1}{2}$	$m_l > l$

**15.** What is the maximum number of electrons that can be accommodated in an n = **7** level?

The sublevels in the n = 7 level: $l =$ 6 5 4 3 2 1 0

The number of orbitals in each sublevel: $(2l+1) =$ 13 + 11 + 9 + 7 + 5 + 3 + 1 = 49

Each orbital can accommodate two electrons so total number of electrons = 2 × 49 = 98 electrons

17. Sketch the p orbitals. Be sure to label the axes.   See Figure 2.11 in text.

19. How many orbitals are in an h sublevel? What is the lowest n quantum number for a level with an h sublevel? How many electrons could occupy this level?

s, p, d, f, g, h → l = 5 for an h sublevel. l+1 = 6, so the n=6 level has an h sublevel. 2 l + 1 = 11 orbitals in an h sublevel, 9 orbitals in a g, 7 in an f, 5 in a d, 3 in a p, and 1 in an s sublevel, for a total of 36 orbitals in the n=6 level. Therefore, the n=6 level could accommodate 72 electrons.

21. Write short hand notations for each of the orbital occupancies shown to the right. Arrows are used to represent the electron spin quantum number. That is, ↑ represents $m_s = +\frac{1}{2}$ and ↓ represents $m_s = -\frac{1}{2}$.

a. $4p^6 5s^2$ b. $2p^6 3s^1$ c. $5s^2 4d^8$
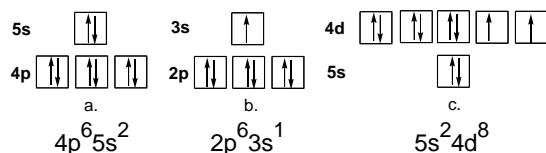
23. Give the n and l quantum numbers for the highest energy electrons in each of the following atoms:

a)	ruthenium	n = 4; l = 2	b)	antimony	n = 5; l = 1
c)	barium	n = 6; l = 0	d)	silicon	n = 3; l = 1

25. Write electron configurations for the following:

a) manganese: $[Ar] 4s^2 3d^5$ b) thallium: $[Xe] 6s^2 4f^{14} 5d^{10} 6p^1$

c) sulfur $[Ne] 3s^2 3p^4$ d) bromine $[Ar] 4s^2 3d^{10} 4p^5$

27. How many of each of the following are populated by at least one electron in an atom of cobalt?

Cobalt is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^7$
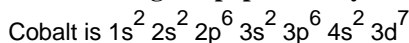
a) levels = 4   b) sublevels = 7   c) orbitals = <u>15</u>

29. How many unpaired electrons are present in each of the following?

a) manganese (5)   b) thallium (1)   c) sulfur   (2)   d) bromine (1)

31. Indicate whether each orbital occupancy shown is a ground state, an excited state or not allowed. If it is a ground state occupancy, identify the atom. If it is not the ground state, explain why.

excited state   excited state   excited state   ground state

orbital energy   Hund's rule   Hund's rule   carbon
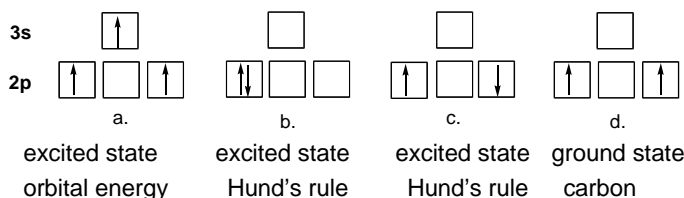
33. What is the common feature of the electron configurations of elements in a given group (family) of the periodic table?

The number of electrons in each of the outermost sublevels is the same.

35. a) Identify the atom that contains five electrons in the n = 3 level.   P

b) Which atom has nine electrons in the n = 3 level?   Sc

37. **Indicate which transition in a hydrogen atom would emit the photon of longer wavelength.**

　　a)　**n = 4 → 3　or　n = 6 → 3**　　　　ans: n = 4 → 3

　　b)　**n = 2 → 1　or　n = 12 → 2**　　　　ans: n = 12 → 2

39. **Give a set of quantum numbers for the electrons in the outermost shell of Ga.**

Outermost shell configuration of Ga = $4s^2 4p^1$

n	l	m_l	m_s
4	0	0	$^1/_2$
4	0	0	$-^1/_2$
4	1	1	$^1/_2$

41. **Draw an orbital energy diagram similar to Figure 2.8 that describes the electrons in the outermost shell of Si and place electrons (arrows) to show the orbital occupancy.**

Silicon's electron configuration is $1s^2 2s^2 2p^6 3s^2 3p^2$, so the outermost shell is the n=3 shell

3p ↑ ↑ □

3s ↑↓
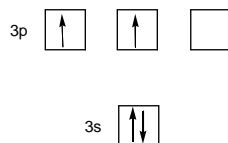
43. **Give the number of s, p, d, and f electrons in each of the following atoms.**

		s	p	d	f
a)	Na	5	6	0	0
b)	Fe	8	12	6	0
c)	Pb	12	26	30	14
d)	Se	8	16	10	0

45. **What are the n and l quantum numbers for the electrons with the highest energy in the following atoms?**

		n	l
a)	Na	3	0
b)	Fe	3	2
c)	Pb	6	1
d)	Se	4	1

47. **Identify the element with each of the following electron configurations:**

　　a)　**$[Ar]4s^2 3d^{10} 4p^3$**　　　　As　　　　　　b)　**$[Ar]4s^2 3d^6$**　　　　Fe

　　c)　**$[Kr]5s^1$**　　　　Rb　　　　　　d)　**$[Xe]6s^2 4f^{14} 5d^{10}$**　　　　Hg

49. **A hydrogen atom in its ground state absorbs a photon of frequency ν = 3.084 x 10^{15} s^{-1}. To what level is the electron promoted?**

Use Equation 2.3a and n_{lo} = 1 for the ground state.

$$3.290 \times 10^{15}(1-\frac{1}{n^2}) = 3.084 \times 10^{15}$$

$$1-\frac{1}{n^2} = \frac{3.084 \times 10^{15}}{3.290 \times 10^{15}} = 0.937$$

$$\frac{1}{n^2} = 1-0.937 = 0.0626 \qquad n^2 = \frac{1}{0.0626} = 16.0 \qquad n = 4$$

The electron is promoted from the n=1 level to the n = 4 level, i.e., it undergoes the 1 → 4 transition.

51. **How many orbitals have each of the following designations?**

　　a) **3p**　　　　　　　　　　　3

　　b) **n = 7 and l =3**　　　　　7

　　c) **n = 3**　　　　　　　　　　9

　　d) **n = 2, l = 1 and m_l = 1**　　　1