Chapter 6

Quantum Theory of the Atom

Chapter 6 suggested problems
10th Ed. and 11th Ed.: 15, 23, 25, 63, 67, 71, 73

links:
Spectra of Gas Discharges at http://astro.u-strasbg.fr/~koppen/discharge/
Atomic Absorption and Emission Spectra at http://csep10.phys.utk.edu/astr162/lect/light/absorption.html
Atomic Spectra at http://hyperphysics.phy-astr.gsu.edu/hbase/quantum/atspect.html

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Class Notes

I. The wave nature of light, quantized energy and photons

A. Radiation

1. Ionizing - comes from radioactive materials (Uranium, plutonium, etc.), consists of atomic nuclei, electrons, or energy

2. Electromagnetic - due to disturbance of electromagnetic fields - of very large items (stars) and very small items (atoms, nuclei)

B. Electromagnetic radiation consists of small, massless packets of energy called photons that travel in waves

1. Amplitude - how intense (high) the wave is

2. Wavelength (λ) - crest to crest distance, can range from kilometers (radiowaves) to picometers (x-rays)

3. Frequency (ν) - how many waves per second (Hz)

4. $c = \lambda \bullet \nu$ where $c = 2.998 \times 10^{10}$ cm/s (the speed of light in a vacuum)

5. The energy of a photon depends on its frequency $E = h\nu = hc/\lambda$

C. Visible light is just a small fraction of the electromagnetic spectrum

1. Radio frequency (30 cm and greater)
2. Microwave (3 mm- 30 cm)
3. Infrared (1000 nm - 3 mm)
4. Visible (400 nm - 800 nm)
5. Ultraviolet (300 nm - 3 nm)
6. X-rays and gamma rays (shorter than 3 nm)

II. Line spectra and the Bohr model

A. 19th century experimentation with cathode-ray tubes and prisms demonstrated that some light sources produce rainbows while other light sources produce a a series of lines

B. Ground state and excited electrons

1. Electrons in their lowest energy state are said to be in the *ground state*, as close to the nucleus as is permitted
2. When an electron absorbs energy (either thermal or electromagnetic) it may become excited and can move to higher unoccupied orbitals, further away from the nucleus

C. Atomic spectra are based on these transitions and are unique for each element (e.g. see Figures 6.11, and 6.12 in the text and links at head of page)

Most common elements in solar spectrum listed in order of decreasing abundance (from "Spectra of Gas Discharges")

Atomic No.	Element	Emission Lines 4000-7000 Å
1	**Hydrogen**	5
2	**Helium**	23
3	**Lithium**	24
8	**Oxygen**	73
6	**Carbon**	27
7	**Nitrogen**	84
10	**Neon**	75
12	**Magnesium**	54
14	**Silicon**	109

16	**Sulfur**	39
26	**Iron**	235
13	**Aluminum**	38
20	**Calcium**	78
18	**Argon**	159
11	**Sodium**	90
36	**Krypton**	75
38	**Strontium**	54
54	**Xenon**	139
56	**Barium**	92

III. Quantum mechanics and quantum numbers

A. In neutral atoms (i.e. atoms in their elemental state) the number of protons is equal to the number of electrons

B. Electrons do not orbit the nucleus in planetary fashion; the position of an electron is described in terms of statistical probability (as calculated using quantum mechanics)

C. Electrons cannot exist just anywhere with respect to the nucleus; they can only be found at certain specific distances from the nucleus (as calculated using quantum mechanics)

D. These distances correspond to energies; the further an electron is from the nucleus the greater its energy

1. $E = h\nu = hc/\lambda$; $h = 6.626 \times 10^{-34}$ J•s

E. Areas of high probability are called *orbitals*

F. The discrete distances at which electrons can be found from the nucleus can be broken down into shells, subshells, and orbitals (see Figures 7.24, 7.25 7.26, 7.27)

1. Principal quantum number (n; n = 1-7)

a. Shell

b. Describes distance from nucleus

c. As n increases, distance increases

d. The spacing between shells is not linear

2. Angular (azimuthal) momentum quantum number (l; l = 0, 1, 2 . . . n-1)

a. Subshell

   b. Each shell contains as many subshells as its number
   c. s, p, d, f, g, h, i
   d. Describes shapes of orbitals
   e. The shape of s, p, d, and f orbitals
   f. The different orbitals are at roughly the same distance from the nucleus but not quite

3. Magnetic quantum number (m_l; -l to +l; 2l +1 total values)
   a. Refer to spatial orientation with respect to Cartesian coordinate system
4. Spin quantum number (m_s; +½ and -½)
   a. Electrons can have one of two spin states - spin up or spin down
5. Any orbital can only hold a maximum of two electrons

shell (n)	subshell (l)	# orbitals	#electrons /orbital	total electrons
1	s	1	2	2
2	s, p	1 + 3 = 4	2	8
3	s, p, d	1 + 3 + 5 = 9	2	18
4	s, p, d, f	1 + 3 + 5 + 7 = 16	2	32
5	s, p, d, f, g	1 + 3 + 5 + 7 + 9 = 25	2	50
6	s, p, d, f, g, h	1 + 3 + 5 + 7 + 9 + 11 = 36	2	72
7	s, p, d, f, g, h, i	1 + 3 + 5 + 7 + 9 + 11 + 13 = 49	2	98

6. g, h, i subshells are virtual and only contain excited electrons
7. The difference in energy between the various energy levels of shells and subshells correspond to specific amounts of energy which in turn correspond to specific wavelengths of light (E = hv)

IV. Electron configurations and the Periodic Table

A. Electron configuration - the distribution of an atom's electrons among available subshells

1. Diagrams are used to show how the orbitals of a subshell are occupied
2. Three types: full configuration, Noble gas core (valence shell), electron arrow diagrams

3. Aufbau Principle (Building-up principle) - start at the bottom and work up
4. Pauli Exclusion Principle: no two electrons can have the same set of four Quantum numbers
5. Hund's rule: all orbitals in a subshell must have at least one electron before any orbital will hold two electrons

B. Note the arrangement of the Periodic Table with respect to electron configuration

1. s, p, d, f blocks
2. Periods and groups give the "address" of the element's highest electron

C. Ground state electron configurations of the first two periods

1. H
2. He
3. Li
4. Mg
5. B
6. C
7. N
8. O
9. F
10. Ne

D. 3rd period elements

E. 4th period elements

F. Exceptions: the chromium and copper groups

G. Noble gas (condensed electron) configurations and electron spin diagrams

H. The electron configuration of ions

V. Inner and outer shell (valence) electrons

A. Inner shell electrons are unreactive under nearly all circumstances, although the electrons can be excited to the point at which they participate in the atomic spectrum of an element

B. Outer shell electrons (valence electrons) are the electrons responsible for chemical reactivity, chemical behavior, bond formation

C. Valence electrons are the reactive outer shell electrons

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