Chapter 8 problems

$c = 299792458\ m\ s^{-1}$ $h = 6.626 \times 10^{-34}\ J\ s$ $N_A = 6.022 \times 10^{23}\ mol^{-1}$

1. What is the frequency in MHz of electromagnetic radiation with a wavelength of 5.6 m?

2. Which have the higher frequency, microwaves or radio waves?
Which have more energy, photons of X-rays or photons of radio waves?

3. Electromagnetic radiation used to heat food in a microwave oven must have a wavelength less than the size of the oven. Assuming that the wavelength is 1×10^{1} cm, what is the frequency of this radiation?

4. Microwave ovens use radiation with a frequency of 2.45 GHz.
a. Calculate the wavelength in cm.
b. What is its energy per photon?

5. Calculate the energy in kJ/mol of photons of electromagnetic radiation that has a frequency of $3.72 \times 10^{17}\ s^{-1}$.

6. How did the study of blackbody radiation and the ultraviolet catastrophe contribute to the development of quantum mechanics?

7. A minimum frequency of 5.56×10^{14} Hz is needed for observation of a photoelectric effect with potassium metal. Light of this frequency is green.
a. With which colors of light will a photoelectric effect be observed for potassium?

b. What is the minimum energy needed to knock an electron out of an atom of potassium metal?

8. What color in the visible spectrum has the lowest energy?

9. What is the significance of the de Broglie equation?

10. What is the wavelength of a neutron moving at a speed of 2200 m/s (zeros are significant)?
$m_{\text{neutron}} = 1.674929 \times 10^{-27}$ kg

11. Explain what happens in an atom when the atom absorbs electromagnetic radiation.

12a. Distinguish between a continuous spectrum and a line spectrum.
b. What conclusion is drawn from the observation that the emission and absorption spectra of atoms are line spectra?

13. What do we mean when we say that something is quantized?
In the Bohr model of the hydrogen atom, what is quantized?

14. For what sizes of particles must one consider quantum effects?
Give an example of such particles.

15. Calculate the wavelength of light emitted when each of the following transitions occur in the hydrogen atom:

$n=3 \rightarrow n=2$

$n=4 \rightarrow n=2$

$n=5 \rightarrow n=2$

$n=6 \rightarrow n=2$

$n=2 \rightarrow n=1$

$n=4 \rightarrow n=3$

$n=\infty \rightarrow n=3$

Use the Balmer-Rydberg formula $\frac{1}{\lambda} = R\left|\frac{1}{n^2} - \frac{1}{n^2}\right|$

where $R = 1.0974 \times 10^{7} m^{-1}$

16. The energy levels for the electron in a hydrogen atom are given by $E = -Rhcn^{-2}$. Make a diagram using $-1/n^2$ in place of the actual energies. Draw energy levels for $n=1,2,3,4,5,6,7,\infty$; show the transitions for problem 15.

17. MULTIPLE CHOICE. Which equation gives the Bohr energy of the electron in a hydrogen atom?

(a) $E = Rhc/n^2$ (b) $E = Rhcn^2$ (c) $E = - Rhc/n^2$ (d) $E = - Rhc/n$ (e) $E = - Rhcn$

18. Define energy.

19. The SI derived unit for energy is J. Rewrite J in terms of SI base units.

20. Give at least two reasons why Bohr's model is no longer used.

21. What is the Heisenberg Uncertainty principle? Write the mathematical relation and explain what it means.

22. What is the physical meaning of a wave function ψ which is a solution of the Schrödinger equation?

23. What does the principal quantum number describe besides energy?

24. Name the four quantum numbers which we use to describe electrons in an atom. What is the significance of each quantum number?

25. In what ways are all *s* orbitals alike?

26. Draw a 2p atomic orbital.

27. When the principal quantum number is 3, what are the possible values for the angular momentum quantum number?

28. When l=1, what are the possible values for m_l?

29. When an orbital is labeled d, the value of l is __.

30. True or False.
__According to the quantum mechanical model, atoms have definite surfaces like the surfaces of space-filling models of atoms.

31. How many electrons can occupy the 3*d* subshell?
How many electrons can be in the third shell?

32. What is meant by effective nuclear charge?

33. Why is Mendeleev rather than Meyer given most of the credit for the periodic table?

34. Draw the orbital box diagram for ...
35. Write the electron configuration for ...

36. What is the importance of the periodic table?

37. What are valence electrons? Why do we emphasize the valence electrons in an atom when discussing atomic properties?

38. Choose the electron configuration that corresponds to an excited state. S&TY8.4
(a) $[Ar]3d^3 4s^2$ (b) $[Kr]5s^1$ (c) $[Ar]3d^5 4s^1$
(d) $[Ne]3s^1 3p^6$ (e) $[Xe]4f^{14} 5d^{10} 6s^2$

39. When the principal quantum number is 4, what are the possible values for the angular momentum quantum number?

+3

40. Distinguish between a 1s and a 2s atomic orbital by size and number of nodes.

+4

41a. When the value of the angular momentum quantum number is 2, how many orbitals are allowed? Explain your answer theoretically.

+6

b. What letter is used to designate these atomic orbitals?

Additional Problems:
Do Practice Examples and Concept Assesments within chapter 8

ANSWER KEY: Chapter 8 problems

$c = 299792458\ m\ s^{-1}$ $h = 6.626 \times 10^{-34}\ J\ s$ $N_A = 6.022 \times 10^{23}\ mol^{-1}$

1. What is the frequency in MHz of electromagnetic radiation with a wavelength of 5.6 m?

$$\nu = \frac{c}{\lambda} = \frac{299792458\ m\ s^{-1}}{5.6\ m} \times \frac{1\ MHz}{10^6\ s^{-1}} = \mathbf{54.\ MHz}$$

2. Which have the higher frequency, **microwaves** or radio waves?
Which have more energy, photons of **X-rays** or photons of radio waves?

3. Electromagnetic radiation used to heat food in a microwave oven must have a wavelength less than the size of the oven. Assuming that the wavelength is 1×10^1 cm, what is the frequency of this radiation?

+4

$$\nu = \frac{c}{\lambda} = \frac{299792458\ m\ s^{-1}}{10\ cm} \times \frac{100\ cm}{1\ m} = \mathbf{3 \times 10^9\ s^{-1}}$$

4. Microwave ovens use radiation with a frequency of 2.45 GHz.
a. Calculate the wavelength in cm.

+4

$$2.45\ GHz \times \frac{10^9\ Hz}{1\ GHz} \times \frac{1\ s^{-1}}{1\ Hz} = 2.45 \times 10^9\ s^{-1}$$

$$\lambda = \frac{c}{\nu} = \frac{2.9979 \times 10^8\ m\ s^{-1}}{2.45 \times 10^9\ s^{-1}} \times \frac{1\ cm}{10^{-2}\ m} = \mathbf{12.2\ cm}$$

b. What is its energy per photon?

+3

$$E = h\nu = (6.626 \times 10^{-34}\ J\ s)(2.45 \times 10^9\ s^{-1}) = \mathbf{1.62 \times 10^{-24}\ J}$$

5. Calculate the energy in kJ/mol of photons of electromagnetic radiation that has a frequency of $3.72 \times 10^{17}\ s^{-1}$. U&B7.10

+4

$$E = h\nu = (6.626 \times 10^{-34}\ J\ s)(3.72 \times 10^{17}\ s^{-1}) \times \frac{1\ kJ}{10^3\ J} \times \frac{6.022 \times 10^{23}}{1\ mol}$$

$$= 1.48 \times 10^5\ kJ/mol$$

6. How did the study of blackbody radiation and the ultraviolet catastrophe contribute to the development of quantum mechanics?

A blackbody is a heated solid object whose emitted radiation does not depend on the identity of the material composing the object. According to classical physics, blackbody radiation should increase in intensity as one goes to shorter and shorter wavelengths, with no intensity maximum, contrary to what is found experimentally. This failure of classical physics at short wavelengths is called the ultraviolet catastrophe. Classical physics assumed that matter could absorb or emit any quantity of energy. Planck found that the observed blackbody radiation profiles could be accounted for by postulating that energy can be gained or lost only in whole-number multiples of hν, implying that energy is quantized. Each small "packet" of energy is called a quantum.

7. A minimum frequency of 5.56×10^{14} Hz is needed for observation of a photoelectric effect with potassium metal. Light of this frequency is green. U&B2 7.57

a. With which colors of light will a photoelectric effect be observed for potassium?

colors with a frequency $\geq 5.56 \times 10^{14}$ Hz
green, blue, and violet

b. What is the minimum energy needed to knock an electron out of an atom of potassium metal?

$$E = h\nu = (6.626 \times 10^{-34}\ \text{J s})(5.56 \times 10^{14}\ \text{s}^{-1}) = 3.68 \times 10^{-19}\ \text{J}$$

8. What color in the visible spectrum has the lowest energy?

red

9. What is the significance of the de Broglie equation?

Matter has wave properties in addition to particle properties.
<u>or</u> **Matter has characteristics of both particles and waves.**

10. What is the wavelength of a neutron moving at a speed of 2200 m/s (zeros are significant)? $m_{\text{neutron}} = 1.674929 \times 10^{-27}$ kg

$$\lambda = \frac{h}{mv} = \frac{6.626 \times 10^{-34}\ \text{J s}}{(1.674929 \times 10^{-27}\ \text{kg})(2200\ \text{m/s})} \times \frac{\text{kg m}^2\ \text{s}^{-2}}{1\ \text{J}}$$

$$= \mathbf{1.798 \times 10^{-10}\ m} \quad \text{or} \quad 0.1798\ \text{nm or} \quad 1.798\ \text{Å}$$

11. Explain what happens in an atom when the atom absorbs electromagnetic radiation. U&B7.34a

When an atom absorbs electromagnetic radiation, an electron gains that energy and moves to a higher energy level.

12a. Distinguish between a continuous spectrum and a line spectrum. U&B7.32

A continuous spectrum contains all wavelengths of electromagnetic radiation over the range that we are observing, such as the range of visible light. A line spectrum contains only a few discrete wavelengths. In the case of the hydrogen atom, the wavelengths are given by the formula

$$\frac{1}{\lambda} = R\left|\frac{1}{n^2} - \frac{1}{n^2}\right|$$

where R = 10 973 731.568 539 m^{-1} and n_{lower} and n_{upper} are integers.

b. What conclusion is drawn from the observation that the emission and absorption spectra of atoms are line spectra?

Only certain energy levels must be available to the electrons in an atom.

13. What do we mean when we say that something is quantized?

it can have only certain discrete (noncontinuous) values

In the Bohr model of the hydrogen atom, what is quantized?

the energy of the electron Z3-7.2

14. For what sizes of particles must one consider quantum effects? Give an example of such particles. Z3-7.3

small masses such as electrons & neutrons

15. Calculate the wavelength of light emitted when each of the following transitions occur in the hydrogen atom:

$n=3 \rightarrow n=2$

$n=4 \rightarrow n=2$

$n=5 \rightarrow n=2$

$n=6 \rightarrow n=2$

$n=2 \rightarrow n=1$

$n=4 \rightarrow n=3$

$n=\infty \rightarrow n=3$

Use the Balmer-Rydberg formula $\frac{1}{\lambda} = R\left|\frac{1}{n^2} - \frac{1}{n^2}\right|$

where $R = 1.0968 \times 10^7 m^{-1}$

n_{lower}	n_{upper}	$1/n_l^2$	$1/n_u^2$	λ (nm)	Color
2	3	0.25	0.111111	656.46	red
2	4	0.25	0.0625	486.26	green
2	5	0.25	0.04	434.16	blue
2	6	0.25	0.027778	410.28	violet
1	2	1	0.25	121.57	UV
3	4	0.11111	0.0625	1875.59	IR
3	10^{151}	0.11111	10^{-302}	820.57	IR
3	∞	0.11111	0	820.57	repeat

16. The energy levels for the electron in a hydrogen atom are given by $E = -Rhcn^{-2}$. Make a diagram using $-1/n^2$ in place of the actual energies. Draw energy levels for $n=1,2,3,4,5,6,7,\infty$; show the transitions for problem 15.

17. MULTIPLE CHOICE. Which equation gives the Bohr energy of the electron in a hydrogen atom?

(a) $E = Rhc/n^2$ (b) $E = Rhcn^2$ **(c) $E = - Rhc/n^2$**

(d) $E = - Rhc/n$ (e) $E = - Rhcn$

18. Define energy.

the capacity to do work or to cause heat flow

19. The SI derived unit for energy is J. Rewrite J in terms of SI base units.

$kg\ m^2\ s^{-2}$

20. Give at least two reasons why Bohr's model is no longer used.
1. Bohr's model only works for the hydrogen atom and ions with one electron outside the nucleus.
2. Bohr's model ascribes definite orbits to electrons, contrary to the Heisenberg Uncertainty principle.

21. What is the Heisenberg Uncertainty principle? Write the mathematical relation and explain what it means.

$\Delta x \cdot \Delta p_x \geq h/4\pi$ where Δ stands for uncertainty, not change
The uncertainty in position times the uncertainty in momentum in the same direction is greater than or equal to Planck's constant divided by 4π. Momentum is mass times velocity.

22. What is the physical meaning of a wave function ψ which is a solution of the Schrödinger equation?
+2 **none**

23. What does the principal quantum number describe besides energy?
+1 **size of the orbital**

24. Name the four quantum numbers which we use to describe electrons in an atom. What is the significance of each quantum number?

Quantum #	Name	Significance
n	principal quantum number	size & energy
l	angular momentum quantum number	shape
m_l	magnetic quantum number	orientation
m_s	electron spin quantum number	spin up or down

25. In what ways are all *s* orbitals alike?
+2 **they are spherical**

26. Draw a 2p atomic orbital.

27. When the principal quantum number is 3, what are the possible values for the angular momentum quantum number?
+3 **2, 1, 0**

28. When l=1, what are the possible values for m_l? **-1, 0, +1**
+3

29. When an orbital is labeled d, the value of l is **2**.
+1

30. True or False.
F According to the quantum mechanical model, atoms have definite surfaces like the surfaces of space-filling models of atoms.

31. How many electrons can occupy the 3*d* subshell?

d ==> $l = 2$ $m_l = -2$ or -1 or 0 or $+1$ or $+2$

Each of these orbitals can hold 2 electrons ==> **10 electrons**

How many electrons can be in the third shell?

$n = 3$	$l = 2$	5 values for m_l	10 electrons
	$l = 1$	3 values for m_l	6 electrons
	$l = 0$	1 value for m_l	2 electrons
			18 total electrons

32. What is meant by effective nuclear charge?

Effective nuclear charge is the apparent nuclear charge exerted on a particular electron, equal to the actual nuclear charge minus the effect of electron repulsions. We often focus on the effective nuclear charge felt by the outermost electron which is shielded significantly by core electrons and somewhat less significantly by other valence electrons.

33. Why is Mendeleev rather than Meyer given most of the credit for the periodic table?

Mendeleev showed how useful the periodic table could be in predicting the existence and properties of yet unknown elements.

34. Draw the orbital box diagram for ...

see pp. 544-546 for first 10 elements

35. Write the electron configuration for ...

see pp. 544-549 for the first 40 elements plus other s- and p-block elements

36. What is the importance of the periodic table?

37. What are valence electrons? Why do we emphasize the valence electrons in an atom when discussing atomic properties? Z3-7.14

outer electrons, or electrons in outermost shell

When atoms interact with each other, the outermost electrons are primarily involved because those electrons are encountered first. The valence electrons are involved in bonding, being shared in covalent bonds and being transferred in ionic bonds. The distance that the valence electrons are from the nucleus helps define the size of an atom.

38. Choose the electron configuration that corresponds to an excited state. S&TY8.4

(a) $[Ar]3d^3 4s^2$ (b) $[Kr]5s^1$ (c) $[Ar]3d^5 4s^1$ **(d) $[Ne]3s^1 3p^6$** (e) $[Xe]4f^{14} 5d^{10} 6s^2$

39. When the principal quantum number is 4, what are the possible values for the angular momentum quantum number?

n-1, n-2, ... , 0 which means the possible values are 3, 2, 1, 0

40. Distinguish between a 1s and a 2s atomic orbital by size and number of nodes.

2s atomic orbital is larger and has one more node

41a. When the value of the angular momentum quantum number is 2, how many orbitals are allowed? Explain your answer theoretically.

$l = 2$ ==> $m_l = -2, -1, 0, +1,$ or $+2$ ==> 5 orbitals

b. What letter is used to designate these atomic orbitals? **d**