Chapter 4 Practice

4.1 The Electromagnetic Spectrum

1. Find the frequency and energy for a photon of light with a wavelength of 375 nm.

   375 nm = 375×10^{-9} m
   $\nu = c/\lambda = (3.00 \times 10^{8}\text{ m/s})(370 \times 10^{-9}\text{ m}) = 8.11 \times 10^{14}$ Hz

2. A photon of light has an energy of 1.83×10^{-18} J. What is the wavelength of this light? Does this light fall in the IR, the visible, or the UV region of the electromagnetic spectrum?

   $E = hc/\lambda$
   $\lambda = hc/E = (6.63 \times 10^{-34}\text{ J·s})(3.00 \times 10^{8}\text{ m/s})/(1.83 \times 10^{-18}\text{ J}) = 1.09 \times 10^{-7}\text{ m} = 109$ nm
   This is shorter than visible light (about 400–700 nm), which means it is higher energy. This light is in the UV region.

4.2 Color, Line Spectra, and the Bohr Model

3. How did the Bohr model describe electrons within an atom? How did this model explain line spectra?
   The Bohr model described electrons orbiting the nucleus, like planets orbit the sun. Only certain orbits were allowed. This model attributed line spectra to light produced when electrons drop from higher- to lower-energy orbits.

4.3 The Quantum Model and Electron Orbitals

4. What is the Heisenberg uncertainty principle? How does this principle affect our description of electrons?
   This principle deals with the mass, velocity, and location of subatomic particles. A central idea of this principle is that it is impossible to know the exact velocity and location of a particle.

5. The Bohr model described electrons in *orbit*. Quantum mechanics describes electrons in *orbitals*. What is the difference between these two terms?
   The Bohr model described electrons as particles orbiting the nucleus, like planets orbit the Sun. Quantum mechanics states that we never actually know the location of the electron. The term *orbital* describes the region around the atom where the electron is most likely to be.

6. How many electrons can occupy each orbital? 2

7. Complete this table, showing the sublevels in each energy level and the number of electrons each can hold.

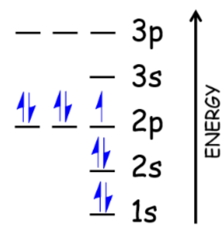
Sublevels (electrons)				*f* (14 electrons)
			d (10 electrons)	*d* (10 electrons)
		p (6 electrons)	*p* (6 electrons)	*p* (6 electrons)
	s (2 electrons)	*s* (2 electrons)	*s* (2 electrons)	*s* (2 electrons)
Energy Level	**1**	**2**	**3**	**4**

4.4 Describing Electron Configurations

8. Complete the sketch below to show the filling sequence for fluorine, then write the electron configuration for this atom.

Fluorine:

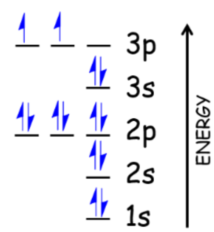
$1s^2 2s^2 2p^5$



9. Complete the sketch below to show the filling sequence for silicon, then write the electron configuration for this atom.

Silicon:

$1s^2 2s^2 2p^6 3s^2 3p^2$



10. Write the electron configurations for each of the following atoms.

N
$1s^2 2s^2 2p^3$

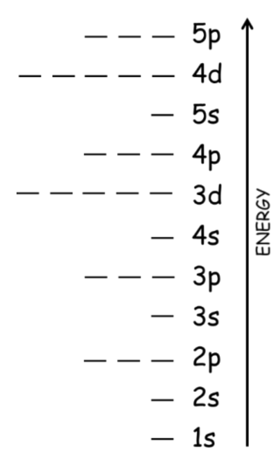
Ne
$1s^2 2s^2 2p^6$

K
$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$

P
$1s^2 2s^2 2p^6 3s^2 3p^3$

Ar
$1s^2 2s^2 2p^6 3s^2 3p^6$

Rb
$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^1$



11. Write electron configurations for the following, using the noble gas shorthand:

Ca
$[Ar]4s^2$

Cd
$[Kr]5s^2 4d^{10}$

Ra
$[Rn]7s^2$

12. Indicate the number of electrons in each ion, then write the electron configuration for each, using the noble gas shorthand.

Br^-	Se^{2-}	Rb^+
35 + 1 = 36 e^- total	34 + 2 = 36 e^- total	37 − 1 = 36 e^- total
[Kr]	[Kr]	[Kr]
Mg^{2+}	Al^{3+}	F^-
12 − 2 = 10 e^- total	13 − 3 = 10 e^- total	9 + 1 = 10 e^- total
[Ne]	[Ne]	[Ne]

4.5 Electron Configuration and the Periodic Table

13. In the table below, label the electron configuration of the outermost shells. The first few are done for you. (Note: For larger atoms, some exceptions to the filling order occur. Don't worry about these—the key idea here is to learn the main pattern.)

This answer shows the main pattern but not the exceptions—the actual filling orders deviate from this slightly.

$1s^1$																	$1s^2$
$2s^1$	$2s^2$											$2p^1$	$2p^2$	$2p^3$	$2p^4$	$2p^5$	$2p^6$
$3s^1$	$3s^2$											$3p^1$	$3p^2$	$3p^3$	$3p^4$	$3p^5$	$3p^6$
$4s^1$	$4s^2$	$3d^1$	$3d^2$	$3d^3$	$3d^4$	$3d^5$	$3d^6$	$3d^7$	$3d^8$	$3d^9$	$3d^{10}$	$4p^1$	$4p^2$	$4p^3$	$4p^4$	$4p^5$	$4p^6$
$5s^1$	$5s^2$	$4d^1$	$4d^2$	$4d^3$	$4d^4$	$4d^5$	$4d^6$	$4d^7$	$4d^8$	$4d^9$	$4d^{10}$	$5p^1$	$5p^2$	$5p^3$	$5p^4$	$5p^5$	$5p^6$
$6s^1$	$6s^2$	$5d^1$	$5d^2$	$5d^3$	$5d^4$	$5d^5$	$5d^6$	$5d^7$	$5d^8$	$5d^9$	$5d^{10}$	$6p^1$	$6p^2$	$6p^3$	$6p^4$	$6p^5$	$6p^6$
$7s^1$	$7s^2$	$6d^1$	$6d^2$	$6d^3$	$6d^4$	$6d^5$	$6d^6$	$6d^7$	$6d^8$	$6d^9$	$6d^{10}$	$7p^1$	$7p^2$	$7p^3$	$7p^4$	$7p^5$	$7p^6$
		$4f^1$	$4f^2$	$4f^3$	$4f^4$	$4f^5$	$4f^6$	$4f^7$	$4f^8$	$4f^9$	$4f^{10}$	$4f^{11}$	$4f^{12}$	$4f^{13}$	$4f^{14}$		
		$5f^1$	$5f^2$	$5f^3$	$5f^4$	$5f^5$	$5f^6$	$5f^7$	$5f^8$	$5f^9$	$5f^{10}$	$5f^{11}$	$5f^{12}$	$5f^{13}$	$5f^{14}$		

14. Which of these is a *d*-block element?	Ca	B	Np	Au
15. Which of these has 7 valence electrons?	Br	O	O^{2-}	Ne
16. Which atom has an electron configuration that is $4s^23d^6$?	Ru	Fe	Ca	P
17. Which of these has an electron configuration of [Ne]?	Na	F	Li^+	Na^+
18. Which two of these have 8 valence electrons?	Br	O	O^{2-}	Ne

19. Which two sublevels form an atom's *valence*?	*s*	*p*	*d*	*f*
20. How many valence electrons do neutral alkaline earth metals have?	1	2	3	4
21. How many valence electrons do noble gases have?	5	6	7	8

22. What is the maximum number of electrons in the highest energy level of an atom? Which sublevels do these occupy?

8 electrons maximum—*s* and *p* sublevels