Chapter 2
The Structure of the Atom and the Periodic Table
Solutions to the Even-Numbered Questions and Problems

In-Chapter Questions and Problems

2.2 a. Phosphorus, atomic number = 15, therefore 15 protons and 15 electrons.
Mass number - atomic number = 30 - 15 = 15, therefore 15 neutrons.
b. Sulfur, atomic number = 16, therefore 16 protons and 16 electrons.
Mass number - atomic number = 32 - 16 = 16, therefore 16 neutrons.
c. Chlorine, atomic number = 17, therefore 17 protons and 17 electrons.
Mass number - atomic number = 35 - 17 = 18, therefore 18 neutrons.

2.4 Bohr described an atomic orbital as a circular path followed by the electron. An orbital is a region of space in the atom, described by a mathematical function, which predicts the probability of finding an electron in the atom.

2.6 a. S, nonmetal
b. O, nonmetal
c. P, nonmetal
d. N, nonmetal

2.8 a. Ar (argon)
b. C (carbon)
c. B (boron)
d. Ar (argon)

2.10 a. magnesium, atomic number = 12, mass = 24.31 amu
b. neon, atomic number = 10, mass = 20.18 amu
c. selenium, atomic number = 34, mass 78.96 amu

2.12 a. Nickel (Ni) has the atomic number 28. Therefore, Ni^{2+} has 28 protons. A neutral nickel atom has 28 electrons and must lose 2 electrons to form the Ni^{2+} cation. As a result, Ni^{2+} has 26 electrons.
b. Bromine (Br) has the atomic number 35. Therefore, Br^{-} has 35 protons. A neutral bromine atom has 35 electrons and must gain 1 more electrons to form the Br^{-} anion. As a result, Br^{-} has 36 electrons.
c. Nitrogen (N) has the atomic number 7. Therefore, N^{3-} has 7 protons. A neutral nitrogen atom has 7 electrons and must gain 3 more electrons to form the N^{3-} anion. As a result, N^{3-} has 10 electrons.

2.14

a.	Rb^{+} has 36 electrons.	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6$	[Kr]
b.	Sr^{2+} has 36 electrons.	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6$	[Kr]
c.	S^{2-} has 18 electrons.	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$	[Ar]
d.	I^{-} has 54 electrons	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^6$	[Xe]

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2.16 a. Cl^{-} has 18 electrons, Ar has 18 electrons; isoelectronic
b. Na^{+} has 10 electrons, Ne has 10 electrons; isoelectronic
c. Mg^{2+} has 10 electrons, Na^{+} has 10 electrons; isoelectronic
d. Li^{+} has 2 electrons, Ne has 10 electrons; not isoelectronic
e. O^{2-} has 10 electrons, F^{-} has 10 electrons; isoelectronic
f. N^{3-} has 10 electrons, Cl- has 18 electrons; not isoelectonic

2.18 a. (smallest) F, Br, I (largest)
b. (lowest) I, Br, F (highest)
c. (lowest) I, Br, F (highest)

End-of-Chapter Questions and Problems

2.20 The mass of the electrons is too small to be significant in comparison to that of the nucleus.

2.22 The isotopic symbol ${}^{13}_{7}C$ is incorrect. The atomic number for carbon is 6. Nitrogen is the element with the atomic number 7.

2.24 a. False. Atoms with a different number of protons are different elements.
b. True
c. False. The mass number of an atom is due to the mass of its protons and neutrons.

2.26 Atoms A and C are isotopes because they both contain the same number of protons but different numbers of neutrons.

2.28 a. atomic number = 17, hence 17 protons and 17 electrons; 37 –17 = 20 neutrons
b. atomic number = 11, hence 11 protons and 11 electrons; 23 –11 = 12 neutrons
c. atomic number = 36, hence 36 protons and 36 electrons; 84 –36 = 48 neutrons

2.30 Atomic number = number of protons = 19, therefore the atom is K (potassium).
Mass number = number of protons + number of neutrons = 19 + 20 = 39
Therefore, the symbol is ${}^{39}_{19}K$

2.32 a. In-115 has an atomic number of 49, so In-115 has 49 protons.
b. In-115 has 115-49 = 66 neutrons.

2.34 a. Iodine has an atomic number of 53, so iodine has 53 protons.
b. Iodine – 131 has 131 – 53 = 78 neutrons.

2.36 a. 1 proton: atomic number = 1, so the element is H
2 neutrons: mass number = 1 + 2 = 3
atomic symbol: ${}^{3}_{1}H$

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b. 92 protons: atomic number = 92, so the element is U
146 neutrons: mass number = 92 + 146 = 238
atomic symbol: ${}^{238}_{92}\mathrm{U}$

2.38 Step 1. Convert each percentage to a decimal fraction.

$$7.49\%\ \text{Lithium-6} \times \frac{1}{100\%} = 0.0749\ \text{Lithium-6}$$

$$92.51\%\ \text{Lithium-7} \times \frac{1}{100\%} = 0.9251\ \text{Lithium-7}$$

Step 2. Multiply the decimal fraction of each isotope by the mass of that isotope to determine the isotopic contribution to the atomic mass.

Contribution to atomic mass by lithium-6	=	fraction of all Li atoms that are lithium-6	X	mass of a lithium-6 atom
	=	0.0749	X	6.0151 amu
	=	0.4505 amu		
Contribution to atomic mass by lithium-7	=	fraction of all Li atoms that are lithium-7	X	mass of a lithium-7 atom
	=	0.9251	X	7.0160 amu
	=	6.4905 amu		

Step 3. The weighted average is:

Atomic mass of naturally occurring Li	=	contribution of lithium-6	+	contribution of lithium-7
	=	0.4505 amu	+	6.4905 amu
	=	6.9410 amu		
	=	6.94 amu (three significant figures)		

2.40 Points of Dalton's theory that are no longer current:

- An atom cannot be created, divided, destroyed, or converted to any other type of atom. (This point was disproved by the discovery of nuclear fusion, fission, and radioactivity.)
- Atoms of a particular element have identical properties. (This point was disproved by the discovery of isotopes.)

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2.42 Crookes used a cathode ray tube. He observed particles emitted by the cathode and traveling toward the anode. This ray was deflected by an electric field. Thomson measured the curvature of the ray influence by the electric and magnetic fields. This measurement provided the mass-to-charge ratio of the negative particle. Thomson also gave the particle the name, electron.

2.44
a. Thomson measured the mass-to-charge ratio of cathode rays and called the negative particle responsible for cathode rays "electrons."
b. Rutherford interpreted the "gold foil experiment" from which he developed the structure of the atom as containing a dense, positive nucleus surrounded by a diffuse sphere of electrons; maintained that the nucleus was responsible for the mass of the atom and the nucleus accounted for only a small fraction of the total volume of the atom.
c. Geiger provided the basic experimental evidence for the existence of the nucleus. A small, dense, positively charged region within the atom was indicated by his alpha-particle scattering experiment.
d. The Bohr theory describes electron arrangement in atoms. Bohr proposed an atomic model that depicted the atom as a nucleus surrounded by fixed energy levels that can be occupied by electrons. He believed that each level was defined by a circular orbit located at some specified distance from the nucleus. Electron promotion from a lower to higher energy level results from absorption of energy that produces an excited state atom. The process of relaxation allows the atom to return to the ground state (the electron falls from a higher to lower energy level) and energy is released.

2.46 Cathode rays are negatively charged. Since opposite charges attract, they deflected toward the positive pole.

2.48 An alpha particle has a positive charge. Since like charges repel, it is deflected away from the positively charged nucleus.

2.50 Rutherford interpreted Geiger's experiment and concluded that the atom is principally empty space. Geiger fired alpha particles at a thin metal foil target and found that most alpha particles passed through the foil instead of being deflected. Rutherford interpreted this to mean that most of the volume of each atom was empty space. If matter was evenly distributed throughout the atom, the matter would have interfered with most of the alpha particles.

2.52 The electromagnetic spectrum is the complete range of electromagnetic waves.

2.54 Electromagnetic radiation, or light, is made up of particles. The energy of each particle is inversely proportional to the wavelength of light.

2.56 The energy of each particle is inversely proportional to the wavelength of light.

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2.58 Infrared radiation has a longer wavelength than ultraviolet radiation. The wavelength of ultraviolet radiation is approximately 10^2 nm, and infrared radiation has wavelenths ranging from 10^3 to 10^5nm.

2.60 Only certain wavelengths that are characteristic of the gas are emitted in the emission spectrum. Different gases have different emission spectra composed of different wavelengths of light.

2.62 The promotion of electrons to higher energy levels requires the absorption of energy.

2.64 Reasons why the Bohr theory did not stand the test of time:
- Although the Bohr theory explains the hydrogen spectrum, it provides only a crude approximation of the spectra for more complex atoms.
- We now believe that electrons could exhibit behavior similar to waves.

2.66 Bohr's model was too simplistic; it confined electrons to narrow regions of space. Consequently, it is limited to simple systems, such as hydrogen.

2.68
a. Calcium (Ca) has an atomic number of 20 and an atomic mass of 40.08.
b. Copper (Cu) has an atomic number of 29 and an atomic mass of 63.55.
c. Cobalt (Co) has an atomic number of 27 and an atomic mass of 58.93.
d. Silicon (Si) has an atomic number of 14 and an atomic mass of 28.09.

2.70 Group IIA or (2) is known collectively as the alkaline earth metals and consists of beryllium (Be), magnesium (Mg), calcium (Ca), strontium (Sr), barium (Ba), and radium (Ra).

2.72 Group VIIIA or (18) is known collectively as the noble gases and consists of helium, neon, argon, krypton, xenon, and radon.

2.74
a. The metals are: Ca, K, Cu, Zn
b. The representative metals are: Ca, K
c. The element that is inert is Kr

2.76 According to the graph prepared for problem 2.65, francium (Fr, atomic number 87) corresponds to a melting point of approximately 26 ºC. The literature value for the melting point of francium is 27 ºC.

2.78 A sublevel is a part of a principal energy level and is designated *s*, *p*, *d*, and *f*. Each sublevel may contain one or more orbitals, regions of space containing a maximum of two electrons with their spins paired.

2.80 A 2s orbital is found in the second principal energy level. A 1s orbital is found in the first principal energy level. A 2s orbital is larger than a 1s orbital.

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2.82 a. two *s* electrons; 2 e^-/orbital x 1 orbital = 2 e^-

b. six *p* electrons; 2 e^-/orbital x 3 orbitals = 6 e^-

c. ten *d* electrons; 2 e^-/orbital x 5 orbitals = 10 e^-

2.84 Hund's rule: When there is a set of orbitals of equal energy, each orbital becomes half-filled before any become completely filled.

a. Orbital diagram does not violate Hund's rule.

b. Orbital diagram does not violate Hund's rule.

c. Orbital diagram violates Hund's rule. Since the three 2p orbitals are of equal energy, two of the orbitals should be half-filled instead of having one 2p orbital completely filled.

2.86 a. Ca $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$

b. Fe $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^6$

c. Cl $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$

2.88 a. V (23 e^-) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^3$

The orbital diagram is:

b. Cd (48 e^-) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}$

The orbital diagram is:

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c. Te (52 e⁻) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^4$

The orbital diagram is:

2.90 a., c., and d. are incorrect.
For a., 4 electrons, the element is helium, and the correct electron configuration is:

$$1s^2\ 2s^2$$

For c., 11 electrons, the element is sodium, and the correct electron configuration is:

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

For d., 5 electrons, the element is boron, and the correct electron configuration is;

$$1s^2\ 2s^2\ 2p^1$$

2.92 The element represented by orbital diagram A is boron. The element represented by orbital diagram B is carbon. The element represented by orbital diagram C is nitrogen.

2.94 a. I has 53 electrons. The noble gas which comes before iodine is Kr. Putting [Kr] in the configuration accounts for the first 36 electrons. The shorthand electron configuration is: [Kr] $5s^2\ 4d^{10}\ 5p^5$

b. Al has 13 electrons. The noble gas which comes before aluminum is Ne. Putting [Ne] in the configuration accounts for the first 10 electrons. The shorthand electron configuration is: [Ne] $3s^2\ 3p^1$

c. V has 23 electrons. The noble gas which comes before vanadium is Ar. Putting [Ar] in the configuration accounts for the first 18 electrons. The shorthand electron configuration is: [Ar] $4s^2\ 3d^3$

2.96 For representative elements, the number of valence electrons in an atom corresponds to the Roman number of the group in which the atom is found.

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2.98 The octet rule can be used to predict the charge of atoms when they become ions. Atoms gain or lose electrons to obtain a noble gas configuration. The ions formed are isoelectronic with the nearest noble gas.

2.100 Nonmetals tend to gain electrons to become negatively charged anions.

2.102 The principal energy level is the same as the period of the periodic table that the element is located in.

	Atom	Total electrons	Valence electrons	Principal energy level number
a.	Mg	12	2	3
b.	K	19	1	4
c.	C	6	4	2
d.	Br	35	7	4
e.	Ar	18	8	3
f.	Xe	54	8	5

2.104 a. Sulfur (S) has the atomic number 16. Therefore, S^{2-} has 16 protons. A neutral sulfur atom has 16 electrons and must gain 2 more electrons to form the S^{2-} anion. As a result, S^{2-} has 18 electrons.
b. Potassium (K) has the atomic number 19. Therefore, K^+ has 19 protons. A neutral potassium atom has 19 electrons and must lose 1 electron to form the K^+ cation. As a result, K^+ has 18 electrons.
b. Cadmium (Cd) has the atomic number 48. Therefore, Cd^{2+} has 48 protons. A neutral cadmium atom has 48 electrons and must lose 2 electrons to form the Cd^{2+} cation. As a result, Cd^{2+} has 46 electrons.

2.106 a. 4 c. 6
b. 5 d. 7

2.108 a. O^{2-} (O gains 2 electrons to attain outermost octet)
b. Br^- (Br gains 1 electron to attain outermost octet)
c. Al^{3+} (Al loses 3 electrons to attain outermost octet)

2.110 a. F^-, 10e⁻; Cl^-, 18 e⁻; Not isoelectronic b. K^+, 18 e⁻; Ar, 18e⁻; Isoelectronic

2.112 a. Ca^{2+} has 18 electrons $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$ [Ar]
b. Mg^{2+} has 10 electrons $1s^2\ 2s^2\ 2p^6$ [Ne]
c. K^+ has 18 electrons $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$ [Ar]
d. Cl^- has 18 electrons $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$ [Ar]

2.114 Atomic size increases from top to bottom down a group in the periodic table.

2.116 Electron affinity is the energy change when a single electron is added to a neutral isolated atom.

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2.118 Energy is released when an electron is added to an isolated chlorine atom.
$Cl + e^- \rightarrow Cl^- + \text{energy}$

2.120 a. (Smallest) P, Si, Al (Largest)
Atomic size decreases as we go across the periodic table within a period.
b. (Smallest) Al, Ga, In (Largest)
Atomic size increases as we go down a group.
c. (Smallest) Ca, Sr, Ba (Largest)
Atomic size increases as we go down a group.
d. (Smallest) N, P, Sb (Largest)
Atomic size increases as we go down a group.

2.122 a. (Smallest) I, Br, Cl (Largest)
Ionization energy generally decreases as we go down a group.
b. (Smallest) Ra, Mg, Be (Largest)
Ionization energy generally decreases as we go down a group.

2.124 a. (Largest) Cl, P, Mg (Smallest)
Electron affinity generally increases as we go across the periodic table within a period.
b. (Largest) Cl, Br, I (Smallest)
Electron affinity generally decreases as we go down a group

2.126 a. A negative ion is always larger than its parent atom because the positive charge in the nucleus is shared among a larger number of electrons in the ion. Each electron moves further from the nucleus and the volume of the ion increases.
b. A sodium ion has a complete octet of electrons and an electron configuration resembling its nearest noble gas. The ionization energy for sodium is quite low; therefore it is easy for an electron to be removed.

2.128 Ar is larger. Each (Ar and K^+) has the same number of electrons (isoelectronic); however, K^+ has one more positive charge in its nucleus, and this excess positive charge draws all electrons closer to the potassium nucleus, making K^+ smaller than the argon atom.

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