2 Atoms, Molecules, and Ions

2.1 It is a metal, and most likely near the end of the transition metals because it can be found in nature in its pure form. A likely candidate is the element marked "b", which is silver.

2.2 (a) K, potassium, metal (b) shiny, metallic solid
(c) It would deform and not crack since metals are malleable.
(d) It would conduct electricity, but it is not a good choice for wiring because it is reacts when exposed to oxygen and humidity (water) in the atmosphere. Potassium is also a soft metal which, makes it unsuitable for use as a wire.

2.3 First, find the S:O ratio in each compound.
Substance A: S:O mass ratio = (6.00 g S) / (5.99 g O) = 1.00
Substance B: S:O mass ratio = (8.60 g S) / (12.88 g O) = 0.668

$$\frac{\text{S:O mass ratio in substance A}}{\text{S:O mass ratio in substance B}} = \frac{1.00}{0.668} = 1.50 = \frac{3}{2}$$

Yes, the mass ratios in the two compounds are simple multiples of each other.

2.4 In compound A the O/S mass ratio is 1. In compound B, the O/S mass ratio is 3/2. This means there is 3/2 times more O in compound B. To find the formula of Compound B multiply the subscript on O in Compound A by 3/2. Compound B is SO_3.

2.5 $$0.005 \text{ mm} \times \frac{1 \times 10^{-3} \text{ m}}{1 \text{ mm}} \times \frac{1 \text{ Au atom}}{2.9 \times 10^{-10} \text{ m}} = 2 \times 10^4 \text{ Au atoms}$$

2.6 $$1 \times 10^{19} \text{ C atoms} \times \frac{1.5 \times 10^{-10} \text{ m}}{\text{C atom}} \times \frac{1 \text{ km}}{1000 \text{ m}} \times \frac{1 \text{ time}}{40{,}075 \text{ km}} = 37.4 \text{ times} \sim 40 \text{ times}$$

2.7 $^{75}_{34}\text{Se}$ has 34 protons, 34 electrons, and (75 – 34) = 41 neutrons.

2.8 The element with 24 protons is Cr. The mass number is the sum of the protons and the neutrons, 24 + 28 = 52. The isotope symbol is $^{52}_{24}\text{Cr}$.

2.9 atomic mass = (62.93)(0.6915) + (64.93)(0.3085) = 63.55
(63.546 from periodic table)

2.10 (a) The atomic weight of Ga is 69.7231. This mass is closer to the mass of gallium-69 than to gallium-71; therefore gallium-69 must be more abundant.
(b) The total abundance of both isotopes must be 100.00%. Let Y be the natural abundance of ^{69}Ga and [1 – Y] the natural abundance of ^{71}Ga.
(68.9256)(Y) + (70.9247)([1 – Y]) = 69.7231

Solve for Y. $Y = \dfrac{-1.2016}{-1.9991} = 0.6011$

^{69}Ga natural abundance = 60.11% and ^{71}Ga natural abundance = 100.00 – 60.11 = 39.89%

2.11 Pt, 195.078

$$\text{mol Pt} = 9.50 \text{ g Pt} \times \frac{1 \text{ mol Pt}}{195.078 \text{ g Pt}} = 0.0487 \text{ mol Pt}$$

$$\text{atoms Pt} = 0.0487 \text{ mol Pt} \times \frac{6.022 \times 10^{23} \text{ atoms Pt}}{1 \text{ mol Pt}} = 2.93 \times 10^{22} \text{ atoms Pt}$$

2.12 $$\text{atomic mass in g} = \frac{1.50 \text{ g}}{2.26 \times 10^{22} \text{ atoms}} \times 6.022 \times 10^{23} \text{ atoms} = 40.0 \text{ g; Y = Ca}$$

2.13 Total intensity = 100.00 + 92.90 = 192.90

mass of isotope 1 = 106.905; mass of isotope 2 = 108.905

$$\text{fractional abundance of isotope 1} = \frac{100.00}{192.90} = 0.5184$$

$$\text{fractional abundance of isotope 2} = \frac{92.90}{192.90} = 0.4816$$

atomic weight = (106.905)(0.5184) + (108.905)(0.4816) = 107.87

The element is Ag.

2.14 Total intensity = 2.672 + 45.992 + 42.176 + 100.000 = 190.840

mass of isotope 1 = 203.9730; mass of isotope 2 = 205.9744

mass of isotope 3 = 206.9758; mass of isotope 4 = 207.9766

$$\text{fractional abundance of isotope 1} = \frac{2.672}{190.840} = 0.01400$$

$$\text{fractional abundance of isotope 2} = \frac{45.992}{190.840} = 0.24100$$

$$\text{fractional abundance of isotope 3} = \frac{42.176}{190.840} = 0.22100$$

$$\text{fractional abundance of isotope 4} = \frac{100.000}{190.840} = 0.524000$$

atomic weight = (203.9730)(0.01400) + (205.9744)(0.24100) + (206.9758)(0.22100) + (207.9766)(0.524000) = 207.217

The element is Pb.

2.15 Figure (b) represents a collection of pure hydrogen peroxide (H_2O_2) molecules.

2.16 (a) Figures (b) and (d) illustrate pure substances.
(b) Figures (a) and (c) illustrate mixtures.
(c) Figures (b) and (d) illustrate the law of multiple proportions.

2.17 thymine, $C_5H_6N_2O_2$

2.18 adrenaline, $C_9H_{13}NO_3$

2.19 (a) LiBr is an ionic compound.

2.20 Figure (a) most likely represents an ionic compound because there are no discrete molecules, only a regular array of two different chemical species (ions). Figure (b) most likely represents a molecular compound because discrete molecules are present.

2.21 (a) magnesium fluoride, MgF_2 (b) tin(IV) oxide, SnO_2

2.22 red – potassium sulfide, K_2S; green – strontium iodide, SrI_2; blue – gallium oxide, Ga_2O_3

2.23 iron(III) carbonate, $Fe_2(CO_3)_3$

2.24 Drawing 1 represents ionic compounds with one cation and two anions. Only (c) $CaCl_2$ is consistent with drawing 1.
Drawing 2 represents ionic compounds with one cation and one anion. Both (a) LiBr and (b) $NaNO_2$ are consistent with drawing 2.

2.25 N_2O_5, dinitrogen pentoxide

2.26 (a) PCl_5, phosphorus pentachloride (b) N_2O, dinitrogen monoxide

2.27 (e)

2.28 (a) false (b) true except for 2 or 3 years prior to 1940 (c) true (d) true

2.29 ^{18}O, 8 protons, 10 neutrons, 8 electrons; ^{2}H, 1 proton, 1 neutron, 1 electron

2.30 Warm seawater near the equator has a higher ratio of $^{18}O/^{16}O$ than snow falling in Antarctica.

2.31 (a) ~14 degrees (b) ~70%

2.32 (a) % ^{18}O = 0.1995% = 0.001 995
% ^{16}O = 100.0000% – 0.1995% = 99.8005% = 0.998 005
seawater O atomic weight = (15.994 914 6)(0.998 005) + (17.999 161 0)(0.001 995)
= 15.998 91 u
(b) % ^{18}O = 0.1971% = 0.001 971
% ^{16}O = 100.0000% – 0.1971% = 99.8029% = 0.998 029
polar ice O atomic weight = (15.994 914 6)(0.998 029) + (17.999 161 0)(0.001 971)
= 15.998 86 u

2.33 H_2O, 18.015 g/mol

(a) $\text{mol } H_2O = 1.00 \times 10^{-6} \text{g } H_2O \times \dfrac{1 \text{ mol } H_2O}{18.015 \text{ mol } H_2O} = 5.55 \times 10^{-8} \text{ mol } H_2O$

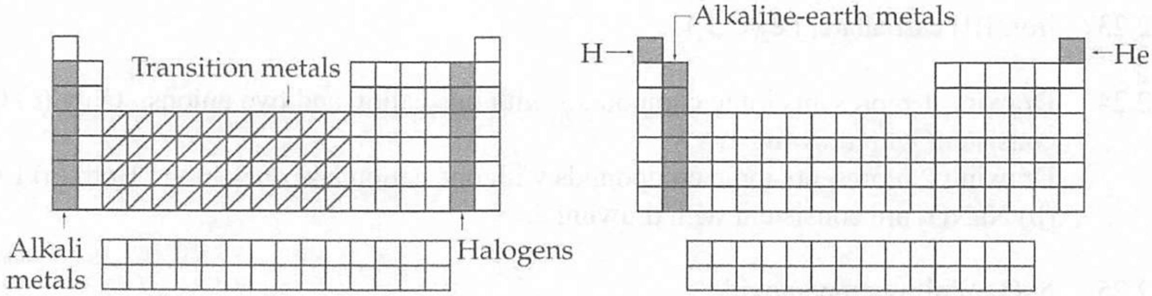
(b) % deuterium = 0.0156% = 0.000 156

$$\text{deuterium atoms} = (5.55 \times 10^{-8} \text{ mol } H_2O)\left(\frac{6.022 \times 10^{23} \ H_2O \text{ molecules}}{1 \text{ mol } H_2O}\right) \times$$

$$\left(\frac{3 \text{ atoms}}{1 \ H_2O \text{ molecule}}\right)\left(\frac{0.000\,156 \text{ D atoms}}{H_2O \text{ atoms}}\right) = 1.56 \times 10^{13} \text{ D atoms}$$

Conceptual Problems

2.34



2.35 red – gas; blue – 42;
green – lithium, sodium, potassium or rubidium are possible answers.

2.36 The element is americium (Am) with atomic number = 95. It is in the actinide series.

2.37 The three "coinage metals" are copper (Cu), silver (Ag), and gold (Au).

2.38 Drawing (a) represents a collection of SO_2 units. Drawing (d) represents a mixture of S atoms and O_2 units.

2.39 To obey the law of mass conservation, the correct drawing must have the same number of red and yellow spheres as in drawing (a). The correct drawing is (d).

2.40 Figures (b) and (c) both contain two protons but different numbers of neutrons. They are isotopes of the same element. Figure (a) contains only one proton. It is a different element than (b) and (c).

2.41 A Na atom has 11 protons and 11 electrons [drawing (b)].
A Ca^{2+} ion has 20 protons and 18 electrons [drawing (c)].
A F^{-} ion has 9 protons and 10 electrons [drawing (a)].

2.42 Figure (b) represents a pure substance consisting of a compound.
Figure (c) represents a pure substance consisting of an element.
Figure (a) represents a mixture of elements.

2.43 methionine, $C_5H_{11}NO_2S$

2.44 (a) alanine, $C_3H_7NO_2$ (b) ethylene glycol, $C_2H_6O_2$ (c) acetic acid, $C_2H_4O_2$

2.45 (a) $MgSO_4$ (b) Li_2CO_3 (c) $FeCl_2$ (d) $Ca_3(PO_4)_2$

Section Problems
Elements and the Periodic Table (Sections 2.1–2.3)

2.46 118 elements are presently known. About 90 elements occur naturally.

2.47 Hydrogen accounts for roughly 75% of the observed mass of the universe. Carbon, hydrogen, nitrogen and oxygen make up 95% of the mass of the human body.

2.48 (a) gadolinium, Gd (b) germanium, Ge (c) technetium, Tc (d) arsenic, As

2.49 (a) cadmium, Cd (b) iridium, Ir (c) beryllium, Be (d) tungsten, W

2.50 (a) Te, tellurium (b) Re, rhenium (c) Be, beryllium (d) Ar, argon
(e) Pu, plutonium

2.51 (a) B, boron (b) Rh, rhodium (c) Cf, californium (d) Os, osmium
(e) Ga, gallium

2.52 (a) Tin is Sn. Ti is titanium. (b) Manganese is Mn. Mg is magnesium.
(c) Potassium is K. Po is polonium.
(d) The symbol for helium is He. The second letter is lowercase.

2.53 (a) The symbol for carbon is C. (b) The symbol for sodium is Na.
(c) The symbol for nitrogen is N. (d) The symbol for chlorine is Cl.

2.54 (a) Mendeleev used the experimentally observed chemistry of the elements to organize them in his table.
(b) Mendeleev realized that there were "holes" in the table below aluminum and silicon. The chemical behavior of aluminum (relative mass ≈ 27.3) is similar to that of boron (relative mass ≈ 11). In the same way, silicon (relative mass ≈ 28) is similar to carbon (relative mass ≈ 12). Using these observations, Mendeleev predicted with remarkable accuracy what the properties of these unknown elements would be. The element immediately below aluminum should have a relative mass near 68 and should have a low melting point. The element below silicon should have a relative mass near 72 and should be dark gray in color.

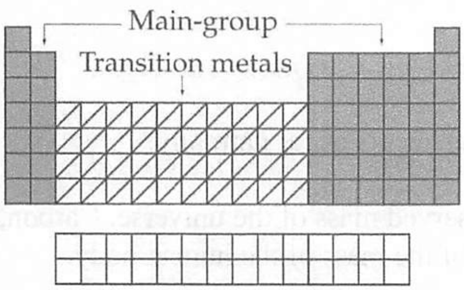
2.55 (a) chemical properties (b) period (c) group (d) Main
(e) Transition metal (f) Mo, metal; Br, nonmetal; Si, semimetal (g) Si

2.56 The rows are called periods, and the columns are called groups.

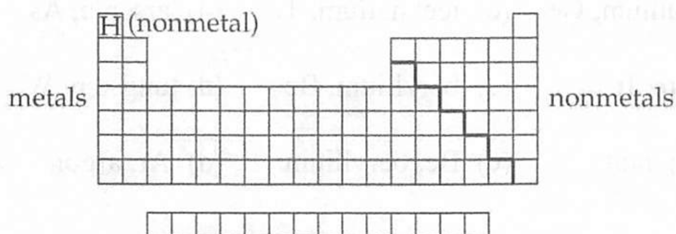
2.57 There are 18 groups in the periodic table. They are labeled as follows:
1A, 2A, 3B, 4B, 5B, 6B, 7B, 8B (3 groups), 1B, 2B, 3A, 4A, 5A, 6A, 7A, 8A

2.58 Elements within a group have similar chemical properties.

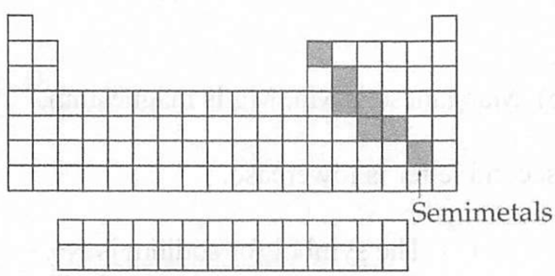
2.59



2.60



2.61



A semimetal is an element with properties that fall between those of metals and nonmetals.

2.62 (a) Ti, metal (b) Te, semimetal (c) Se, nonmetal
(d) Sc, metal (e) Si, semimetal

2.63 (a) Ar, nonmetal (b) Sb, semimetal (c) Mo, metal
(d) Cl, nonmetal (e) N, nonmetal (e) Mg, metal

2.64 (a) The alkali metals are shiny, soft, low-melting metals that react rapidly with water to form products that are alkaline.
(b) The noble gases are gases of very low reactivity.
(c) The halogens are nonmetallic and corrosive. They are found in nature only in combination with other elements.

2.65 (a) Li, Na, K, Rb, and Cs (b) Be, Mg, Ca, Sr, and Ba

2.66 F, Cl, Br, and I

2.67 He, Ne, Ar, Kr, Xe, and Rn

2.68 An element that is a soft, silver-colored solid that reacts violently with water and is a good conductor of electricity has the characteristics of a metal.

2.69 An element that is a brittle, shiny, silver-colored solid that is a poor conductor of electricity has the characteristics of a semimetal.

2.70 An element that is a yellow crystalline solid, does not conduct electricity, and when hit with a hammer it shatters has the characteristics of a nonmetal.

2.71 An element that is a colorless, unreactive gas has the characteristics of a nonmetal.

2.72 All match in groups 2A and 7A.

2.73 In group 1A, sodium (Na) and potassium (K), and in group 5A, antimony (Sb).

2.74 Intensive properties do not depend on the amount of substance present.

2.75 Volume and mass are extensive properties. Density and electrical conductivity are intensive properties.

Atomic Theory (Sections 2.4–2.5)

2.76 The law of mass conservation in terms of Dalton's atomic theory states that chemical reactions only rearrange the way that atoms are combined; the atoms themselves are not changed.
The law of definite proportions in terms of Dalton's atomic theory states that the chemical combination of elements to make different substances occurs when atoms join together in small, whole-number ratios.

2.77 The law of multiple proportions states that if two elements combine in different ways to form different substances, the mass ratios are small, whole-number multiples of each other. This is very similar to Dalton's statement that the chemical combination of elements to make different substances occurs when atoms join together in small, whole-number ratios.

2.78 In any chemical reaction, the combined mass of the final products equals the combined mass of the starting reactants.
mass of reactants = mass of products
mass of reactants = mass of Hg + mass of O_2 = 114.0 g + 12.8 g = 126.8 g
mass of products = mass of HgO + mass of left over O_2 = 123.1 g + mass of left over O_2

126.8 g = 123.1 g + mass of left over O_2
mass of left over O_2 = 126.8 g – 123.1 g = 3.7 g

2.79 In any chemical reaction, the combined mass of the final products equals the combined mass of the starting reactants.
mass of reactants = mass of products
mass of reactants = mass of $CaCO_3$ = 612 g
mass of products = mass of CaO + mass of CO_2 = 343 g + mass of CO_2
612 g = 343 g + mass of CO_2
mass of CO_2 = 612 g – 343 g = 269 g

2.80 For the "other" compound: C:H mass ratio = (32.0 g C) / (8.0 g H) = 4
The "other" compound is not methane because the methane C:H mass ratio is 3.

$$\frac{\text{C:H mass ratio in "other"}}{\text{C:H mass ratio in methane}} = \frac{4}{3}$$

2.81 For the "other" compound: B:H mass ratio = (43.2 g B) / (6.0 g H) = 7.2
The "other" compound is not borane because the borane B:H mass ratio is 3.6.

$$\frac{\text{B:H mass ratio in "other"}}{\text{B:H mass ratio in borane}} = \frac{7.2}{3.6} = \frac{2}{1}$$

2.82 First, find the C:H ratio in each compound.
Benzene: C:H mass ratio = (4.61 g C) / (0.39 g H) = 12
Ethane: C:H mass ratio = (4.00 g C) / (1.00 g H) = 4.00
Ethylene: C:H mass ratio = (4.29 g C) / (0.71 g H) = 6.0

$$\frac{\text{C:H mass ratio in benzene}}{\text{C:H mass ratio in ethane}} = \frac{12}{4.00} = \frac{3}{1}$$

$$\frac{\text{C:H mass ratio in benzene}}{\text{C:H mass ratio in ethylene}} = \frac{12}{6.0} = \frac{2}{1}$$

$$\frac{\text{C:H mass ratio in ethylene}}{\text{C:H mass ratio in ethane}} = \frac{6.0}{4.00} = \frac{3}{2}$$

2.83 (a) For benzene:

$$4.61 \text{ g} \times \frac{1 \text{ u}}{1.6605 \times 10^{-24} \text{ g}} \times \frac{1 \text{ C atom}}{12.011} = 2.31 \times 10^{23} \text{ C atoms}$$

$$0.39 \text{ g} \times \frac{1 \text{ u}}{1.6605 \times 10^{-24} \text{ g}} \times \frac{1 \text{ H atom}}{1.008} = 2.3 \times 10^{23} \text{ H atoms}$$

$$\frac{\text{C}}{\text{H}} = \frac{2.31 \times 10^{23} \text{ C atoms}}{2.3 \times 10^{23} \text{ H atoms}} = \frac{1 \text{ C}}{1 \text{ H}}$$

A possible formula for benzene is CH.

For ethane:

$$4.00 \text{ g} \times \frac{1 \text{ u}}{1.6605 \times 10^{-24} \text{ g}} \times \frac{1 \text{ C atom}}{12.011} = 2.01 \times 10^{23} \text{ C atoms}$$

$$1.00 \text{ g} \times \frac{1 \text{ u}}{1.6605 \times 10^{-24} \text{ g}} \times \frac{1 \text{ H atom}}{1.008} = 5.97 \times 10^{23} \text{ H atoms}$$

$$\frac{\text{C}}{\text{H}} = \frac{2.01 \times 10^{23} \text{ C atoms}}{5.97 \times 10^{23} \text{ H atoms}} = \frac{1 \text{ C}}{3 \text{ H}}$$ A possible formula for ethane is CH_3.

For ethylene:

$$4.29 \text{ g} \times \frac{1 \text{ u}}{1.6605 \times 10^{-24} \text{ g}} \times \frac{1 \text{ C atom}}{12.011} = 2.15 \times 10^{23} \text{ C atoms}$$

$$0.71 \text{ g} \times \frac{1 \text{ u}}{1.6605 \times 10^{-24} \text{ g}} \times \frac{1 \text{ H atom}}{1.008} = 4.2 \times 10^{23} \text{ H atoms}$$

$$\frac{\text{C}}{\text{H}} = \frac{2.15 \times 10^{23} \text{ C atoms}}{4.2 \times 10^{23} \text{ H atoms}} = \frac{1 \text{ C}}{2 \text{ H}}$$ A possible formula for ethylene is CH_2.

(b) The results in part (a) give the smallest whole-number ratio of C to H for benzene ($CH \times 6 = C_6H_6$), ethane ($CH_3 \times 2 = C_2H_6$), and ethylene ($CH_2 \times 2 = C_2H_4$), and these ratios are consistent with their modern formulas.

2.84 Assume a 100.0 g sample for each compound and then find the O:C ratio in each compound.
Compound 1: O:C mass ratio = (57.1 g O)/(42.9 g C) = 1.33
Compound 2: O:C mass ratio = (72.7 g O)/(27.3 g C) = 2.66

$$\frac{\text{O:C mass ratio in compound 2}}{\text{O:C mass ratio in compound 1}} = \frac{2.66}{1.33} = \frac{2}{1}$$

If compound 1 is CO and the O:C mass ratio is 2 times that of compound 1, then the formula for compound 2 is CO_2.

2.85 First, find the O:C ratio in each compound.
Carbon suboxide: O:C mass ratio = (1.18 g O)/(1.32 g C) = 0.894
Carbon monoxide: O:C mass ratio = (16.00 g O)/(12.00 g C) = 1.33

$$\frac{\text{O:C mass ratio for carbon suboxide}}{\text{O:C mass ratio for CO}} = \frac{0.894}{1.33} = \frac{0.67}{1}$$

The O:C mass ratio for carbon suboxide is 2/3 that for CO, then the formula for carbon suboxide is $CO_{2/3}$. To get integer subscripts, multiply both subscripts by 3. The formula for carbon suboxide is C_3O_2.

Elements and Atoms (Sections 2.6–2.8)

2.86 electron

2.87 The strength of the magnetic or electric field affects the magnitude of the deflection of the cathode ray in Thomson's experiment.

2.88 (a) true (b) true (c) true (d) true (e) false (f) true

2.89 mass, electrons, negative, positive, a whole number multiple, charge, mass

2.90 (a) -1.010×10^{-18} C because it is not an integer multiple of the electron charge.

2.91 The results of Rutherford's gold foil experiment showed that atoms contained a very small but massive positively charged nucleus.

2.92 (a) The alpha particles would pass right through the gold foil with little to no deflection.

2.93 Assume the nucleus has a diameter of 1 mm.
$1 \text{ mm} = 1 \times 10^{-3} \text{ m}$

$$1 \times 10^{-3} \text{ m (nucleus diameter)} \times \frac{1 \times 10^{-10} \text{ m (atom diameter)}}{1 \times 10^{-15} \text{ m (nucleus diameter)}} = 100 \text{ m (atom diameter)}$$

2.94 $350 \text{ pm} = 350 \times 10^{-12} \text{ m}$

$$\text{Pb atoms} = 0.25 \text{ in} \times \frac{2.54 \text{ cm}}{1 \text{ in}} \times \frac{1 \times 10^{-2} \text{ m}}{1 \text{ cm}} \times \frac{1 \text{ Pb atom}}{350 \times 10^{-12} \text{ m}} = 1.8 \times 10^{7} \text{ Pb atoms thick}$$

2.95 $1 \text{ mm} = 1 \times 10^{-3} \text{ m}$; $150 \text{ pm} = 150 \times 10^{-12} \text{ m}$

$$\text{C atoms} = 1 \times 10^{-3} \text{ m} \times \frac{1 \text{ C atom}}{150 \times 10^{-12} \text{ m}} = 7 \times 10^{6} \text{ C atoms}$$

2.96 The atomic number is equal to the number of protons.
The mass number is equal to the sum of the number of protons and the number of neutrons.

2.97 The atomic number is equal to the number of protons.
The atomic mass is the weighted average mass of the various isotopes for a particular element.

2.98 The subscript giving the atomic number of an atom is often left off an isotope symbol because one can readily look up the atomic number in the periodic table.

2.99 Te has isotopes with more neutrons than the isotopes of I.

2.100 (a) carbon, C (b) argon, Ar (c) vanadium, V

2.101 $^{137}_{55}\text{Cs}$

2.102 (a) $^{220}_{86}\text{Rn}$ (b) $^{210}_{84}\text{Po}$ (c) $^{197}_{79}\text{Au}$

2.103 (a) $^{140}_{58}\text{Ce}$ (b) $^{60}_{27}\text{Co}$

2.104 (a) $^{15}_{7}\text{N}$, 7 protons, 7 electrons, (15 – 7) = 8 neutrons
(b) $^{60}_{27}\text{Co}$, 27 protons, 27 electrons, (60 – 27) = 33 neutrons

(c) $^{131}_{53}I$, 53 protons, 53 electrons, (131 – 53) = 78 neutrons

(d) $^{142}_{58}Ce$, 58 protons, 58 electrons, (142 – 58) = 84 neutrons

2.105 (a) ^{27}Al, 13 protons and (27 – 13) = 14 neutrons

(b) ^{32}S, 16 protons and (32 – 16) = 16 neutrons

(c) ^{64}Zn, 30 protons and (64 – 30) = 34 neutrons

(d) ^{207}Pb, 82 protons and (207 – 82) = 125 neutrons

2.106 (a) $^{24}_{12}Mg$, magnesium (b) $^{58}_{28}Ni$, nickel

(c) $^{104}_{46}Pd$, palladium (d) $^{183}_{74}W$, tungsten

2.107 (a) $^{202}_{80}Hg$, mercury (b) $^{195}_{78}Pt$, platinum

(c) $^{184}_{76}Os$, osmium (d) $^{209}_{83}Bi$, bismuth

2.108 $^{12}_{5}C$, the atomic number for carbon is 6, not 5.

$^{33}_{35}Br$, the mass number must be greater than the atomic number.

$^{11}_{5}Bo$, the element symbol for boron is B.

2.109 $^{14}_{7}Ni$, the element symbol for nitrogen is N.

$^{73}_{23}Ge$, the atomic number for germanium is 32, not 23.

$^{1}_{2}He$, the mass number must be greater than the atomic number.

2.110 Deuterium is ^{2}H and deuterium fluoride is ^{2}HF.
^{2}H has 1 proton, 1 neutron, and 1 electron.
F has 9 protons, 10 neutrons, and 9 electrons.
^{2}HF has 10 protons, 11 neutrons, and 10 electrons.
Chemically, ^{2}HF is like HF and is a weak acid.

2.111 $^{1}H^{35}Cl$ has 18 protons, 18 neutrons, and 18 electrons.
$^{1}H^{37}Cl$ has 18 protons, 20 neutrons, and 18 electrons.
$^{2}H^{35}Cl$ has 18 protons, 19 neutrons, and 18 electrons.
$^{2}H^{37}Cl$ has 18 protons, 21 neutrons, and 18 electrons.
$^{3}H^{35}Cl$ has 18 protons, 20 neutrons, and 18 electrons.
$^{3}H^{37}Cl$ has 18 protons, 22 neutrons, and 18 electrons.

Atomic Weight, Moles, and Mass Spectrometry (Sections 2.9–2.10)

2.112 (b) ^{12}C

2.113 (a) 1 u = 1.660 539 x 10^{-24} g (b) 1 mol of atoms = 6.022 141 x 10^{23} atoms

2.114 (a) atomic mass (b) atomic number (c) molar mass
(d) mass number (e) atomic weight

2.115 (a) false (b) true

2.116 An element's atomic mass is the weighted average of the isotopic masses of the element's naturally occurring isotopes. The atomic mass for Cu (63.546) must fall between the masses of its two isotopes. If one isotope is ^{65}Cu, the other isotope must be ^{63}Cu, and not ^{66}Cu. If the other isotope was ^{66}Cu, the atomic mass for Cu would be greater than 65.

2.117 An element's atomic mass is the weighted average of the isotopic masses of the element's naturally occurring isotopes. The atomic mass for S (32.065) must fall between the masses of its lightest and heaviest isotopes. If three of its isotopes are ^{33}S, ^{34}S, and ^{36}S, the other isotope must be ^{32}S, and not ^{35}S. If the other isotope was ^{35}S, the atomic mass for S would be greater than 33.

2.118 (10.0129)(0.199) + (11.009 31)(0.801) = 10.8 for B

2.119 (106.9051)(0.5184) + (108.9048)(0.4816) = 107.9 for Ag

2.120 24.305 = (23.985)(0.7899) + (24.986)(0.1000) + (Z)(0.1101)
Solve for Z. Z = 25.982 for ^{26}Mg.

2.121 The total abundance of all three isotopes must be 100.00%. The natural abundance of ^{29}Si is 4.68%. The natural abundance of ^{28}Si and ^{30}Si together must be 100.00% – 4.68% = 95.32%. Let Y be the natural abundance of ^{28}Si and [95.32 – Y] the natural abundance of ^{30}Si.
28.0855 = (28.9765)(0.0468) + (27.9769)(Y) + (29.9738)([0.9532 – Y])

Solve for Y. $Y = \frac{-1.842}{-1.997} = 0.9224$

^{28}Si natural abundance = 92.2% ^{30}Si natural abundance = 95.32 – 92.24 = 3.1%

2.122 atomic weight = (62.93)(0.6915) + (64.93)(0.3085) = 63.55

2.123 atomic weight = (69.924)(0.205) + (71.922)(0.274) + (72.923)(0.078)
+ (73.921)(0.365) + (75.921)(0.078) = 72.6

2.124 (a) $\text{g Ti} = 1.505 \text{ mol Ti} \times \dfrac{47.867 \text{ g Ti}}{1 \text{ mol Ti}} = 72.04 \text{ g Ti}$

(b) $\text{g Na} = 0.337 \text{ mol Na} \times \dfrac{22.989\ 770 \text{ g Na}}{1 \text{ mol Na}} = 7.75 \text{ g Na}$

(c) $\text{g U} = 2.583 \text{ mol U} \times \dfrac{238.028\ 91 \text{ g U}}{1 \text{ mol U}} = 614.8 \text{ g U}$

2.125 (a) $\text{mol Ti} = 11.51 \text{ g Ti} \times \dfrac{1 \text{ mol Ti}}{47.867 \text{ g Ti}} = 0.2405 \text{ mol Ti}$

(b) $\text{mol Na} = 29.127 \text{ g Na} \times \dfrac{1 \text{ mol Na}}{22.989\ 770 \text{ g Na}} = 1.2670 \text{ mol Na}$

(c) $\text{mol U} = 1.477 \text{ kg} \times \dfrac{1000 \text{ g}}{1 \text{ kg}} \times \dfrac{1 \text{ mol U}}{238.028\ 91 \text{ g U}} = 6.205 \text{ mol U}$

2.126 The mass of 6.02×10^{23} atoms is its atomic mass expressed in grams. If the atomic mass of an element is X, then 6.02×10^{23} atoms of this element weighs X grams.

2.127 $\text{mass} = \dfrac{x \text{ g}}{6.02 \times 10^{23} \text{ atoms}} \times 3.17 \times 10^{20} \text{ atoms} = (x) \times 5.27 \times 10^{-4} \text{ g}$

2.128 The mass of 6.02×10^{23} atoms is its atomic mass expressed in grams. If the mass of 6.02×10^{23} atoms of element Y is 83.80 g, then the atomic mass of Y is 83.80. Y is Kr.

2.129 $\text{atomic mass in g} = \dfrac{0.815 \text{ g}}{4.61 \times 10^{21} \text{ atoms}} \times 6.02 \times 10^{23} \text{ atoms} = 106 \text{ g}; \ Z = \text{Pd}$

2.130 (a) The purpose of bombarding gaseous atoms with an electron beam is to knock electrons off of the atoms, leaving them with a positive charge.
(b) A lighter ion will be deflected to a greater degree by the magnetic field.
(c) The magnetic field strength is adjusted so that a different mass-to-charge ratio ion strikes the detector.

2.131 (a) Total intensity = 50 + 100 = 150

fractional abundance of $^{63}\text{Cu} = \dfrac{50}{150} = 0.33$

fractional abundance of $^{65}\text{Cu} = \dfrac{100}{150} = 0.67$

(b) Total intensity = 100 + 45 = 145

fractional abundance of $^{63}\text{Cu} = \dfrac{100}{145} = 0.69$

fractional abundance of $^{65}Cu = \dfrac{45}{145} = 0.31$

(c) Total intensity = 100 + 70 = 170

fractional abundance of $^{63}Cu = \dfrac{100}{170} = 0.59$

fractional abundance of $^{65}Cu = \dfrac{70}{170} = 0.41$

Spectrum (b) represents a sample of Cu because it yields the correct fractional abundances.

2.132 Total intensity = 100.00 + 74.80 = 174.80
mass of isotope 1 = 120.9038; mass of isotope 2 = 122.9042

fractional abundance of isotope 1 = $\dfrac{100.00}{174.80} = 0.5721$

fractional abundance of isotope 2 = $\dfrac{74.80}{174.80} = 0.4279$

atomic weight = (120.9038)(0.5721) + (122.9042)(0.4279) = 121.76
The element is Sb.

2.133 Total intensity = 11.39 + 10.09 + 100.00 + 7.43 + 7.03 = 135.94
mass of isotope 1 = 45.9527; mass of isotope 2 = 46.9518
mass of isotope 3 = 47.9479; mass of isotope 4 = 48.9479
mass of isotope 5 = 49.9448

fractional abundance of isotope 1 = $\dfrac{11.39}{135.94} = 0.08379$

fractional abundance of isotope 2 = $\dfrac{10.09}{135.94} = 0.07422$

fractional abundance of isotope 3 = $\dfrac{100.00}{135.94} = 0.7356$

fractional abundance of isotope 4 = $\dfrac{7.43}{135.94} = 0.0547$

fractional abundance of isotope 5 = $\dfrac{7.03}{135.94} = 0.0517$

atomic weight = (45.9527)(0.083 79) + (46.9518)(0.074 22) + (47.9479)(0.7356) + (48.9479)(0.0547) +(49.9448)(0.0517) = 47.87
The element is Ti.

Chemical Compounds (Sections 2.11–2.12)

2.134 A covalent bond results when two atoms share several (usually two) of their electrons. An ionic bond results from a complete transfer of one or more electrons from one atom to another. The C–H bonds in methane (CH_4) are covalent bonds. The bond in NaCl (Na^+Cl^-) is an ionic bond.

2.135 Covalent bonds typically form between nonmetals. (a) B–Br, (c) Br–Cl, and (d) O–Br are covalent bonds.
Ionic bonds typically form between a metal and a nonmetal. (b) Na–Br is an ionic bond.

2.136 Element symbols are composed of one or two letters. If the element symbol is two letters, the first letter is uppercase and the second is lowercase. CO stands for carbon and oxygen in carbon monoxide.

2.137 (a) The formula of ammonia is NH_3.
(b) The ionic solid potassium chloride has the formula KCl.
(c) Cl^- is an anion.
(d) CH_4 is a neutral molecule.

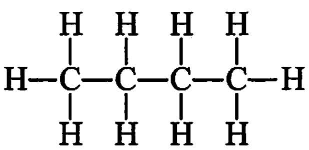
2.138 (a) Be^{2+}, 4 protons and 2 electrons
(b) Rb^+, 37 protons and 36 electrons
(c) Se^{2-}, 34 protons and 36 electrons
(d) Au^{3+}, 79 protons and 76 electrons

2.139 (a) A +2 cation that has 36 electrons must have 38 protons. X = Sr.
(b) A –1 anion that has 36 electrons must have 35 protons. X = Br.

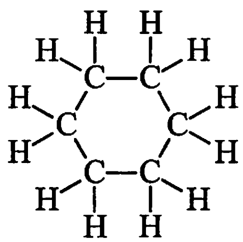
2.140 C_3H_8O

2.141 $C_3H_6O_3$

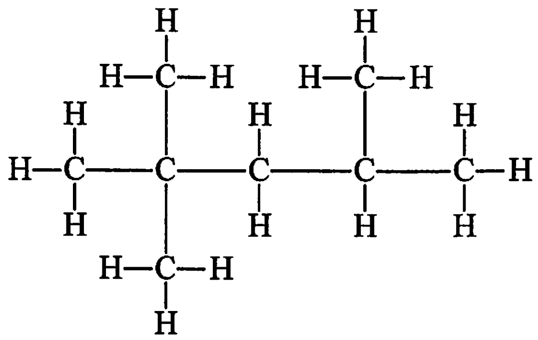
2.142



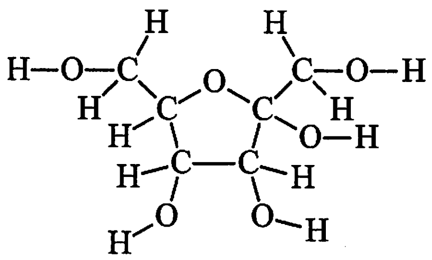
2.143



2.144



2.145



Naming Compounds (Section 2.13)

2.146 (a) CsF, cesium fluoride (b) K_2O, potassium oxide (c) CuO, copper(II) oxide

2.147 (a) BaS, barium sulfide (b) $BeBr_2$, beryllium bromide (c) $FeCl_3$, iron(III) chloride

2.148 (a) potassium chloride, KCl (b) tin(II) bromide, $SnBr_2$ (c) calcium oxide, CaO
(d) barium chloride, $BaCl_2$ (e) aluminum hydride, AlH_3

2.149 (a) vanadium(III) chloride, VCl_3 (b) manganese(IV) oxide, MnO_2
(c) copper(II) sulfide, CuS (d) aluminum oxide, Al_2O_3

2.150 (a) calcium acetate, $Ca(CH_3CO_2)_2$ (b) iron(II) cyanide, $Fe(CN)_2$
(c) sodium dichromate, $Na_2Cr_2O_7$ (d) chromium(III) sulfate, $Cr_2(SO_4)_3$
(e) mercury(II) perchlorate, $Hg(ClO_4)_2$

2.151 (a) lithium phosphate, Li_3PO_4 (b) magnesium hydrogen sulfate, $Mg(HSO_4)_2$
(c) manganese(II) nitrate, $Mn(NO_3)_2$ (d) chromium(III) sulfate, $Cr_2(SO_4)_3$

2.152 (a) $Ca(ClO)_2$, calcium hypochlorite
(b) $Ag_2S_2O_3$, silver(I) thiosulfate or silver thiosulfate
(c) NaH_2PO_4, sodium dihydrogen phosphate (d) $Sn(NO_3)_2$, tin(II) nitrate
(e) $Pb(CH_3CO_2)_4$, lead(IV) acetate (f) $(NH_4)_2SO_4$, ammonium sulfate

2.153 (a) Ba^{2+}, barium ion (b) Cs^+, cesium ion (c) V^{3+}, vanadium(III) ion
(d) HCO_3^-, hydrogen carbonate ion (e) NH_4^+, ammonium ion (f) Ni^{2+}, nickel(II) ion
(g) NO_2^-, nitrite ion (h) ClO_2^-, chlorite ion
(i) Mn^{2+}, manganese(II) ion (j) ClO_4^-, perchlorate ion

2.154 (a) $CaBr_2$ (b) $CaSO_4$ (c) $Al_2(SO_4)_3$

2.155 (a) $NaNO_3$ (b) K_2SO_4 (c) $SrCl_2$

2.156 (a) $CaCl_2$ (b) CaO (c) CaS

2.157 (a) RbBr (b) Rb_3N (c) Rb_2Se

2.158 (a) sulfite ion, SO_3^{2-} (b) phosphate ion, PO_4^{3-} (c) zirconium(IV) ion, Zr^{4+}
(d) chromate ion, CrO_4^{2-} (e) acetate ion, $CH_3CO_2^-$ (f) thiosulfate ion, $S_2O_3^{2-}$

2.159 (a) Zn^{2+} (b) Fe^{3+} (c) Ti^{4+} (d) Sn^{2+} (e) Hg_2^{2+} (f) Mn^{4+} (g) K^+ (h) Cu^{2+}

2.160 (a) CCl_4, carbon tetrachloride (b) ClO_2, chlorine dioxide
(c) N_2O, dinitrogen monoxide (d) N_2O_3, dinitrogen trioxide

2.161 (a) NCl_3, nitrogen trichloride (b) P_4O_6, tetraphosphorus hexoxide
(c) S_2F_2, disulfur difluoride

2.162 (a) NO, nitrogen monoxide (b) N_2O, dinitrogen monoxide (c) NO_2, nitrogen dioxide
(d) N_2O_4, dinitrogen tetroxide (e) N_2O_5, dinitrogen pentoxide

2.163 (a) SO, sulfur monoxide (b) SO_2, disulfur dioxide (c) S_5O, pentasulfur monoxide
(d) S_7O_2, heptasulfur dioxide (e) SO_3, sulfur trioxide

2.164 (a) Na^+ and SO_4^{2-}; therefore the formula is Na_2SO_4
(b) Ba^{2+} and PO_4^{3-}; therefore the formula is $Ba_3(PO_4)_2$
(c) Ga^{3+} and SO_4^{2-}; therefore the formula is $Ga_2(SO_4)_3$

2.165 (a) sodium peroxide, Na_2O_2 (b) aluminum bromide, $AlBr_3$
(c) chromium(III) sulfate, $Cr_2(SO_4)_3$

Multiconcept Problems

2.166 For NH_3, $$(2.34\text{ g N})\left(\frac{3 \times 1.0079\text{ H}}{14.0067\text{ N}}\right) = 0.505\text{ g H}$$

For N_2H_4, $$(2.34\text{ g N})\left(\frac{4 \times 1.0079\text{ H}}{2 \times 14.0067\text{ N}}\right) = 0.337\text{ g H}$$

2.167 $$\text{g N} = (1.575\text{ g H})\left(\frac{3.670\text{ g N}}{0.5275\text{ g H}}\right) = 10.96\text{ g N}$$

From Problem 2.166:

for NH_3, $\frac{g\,N}{g\,H} = \frac{2.34\,g\,N}{0.505\,g\,H} = 4.63$

for N_2H_4, $\frac{g\,N}{g\,H} = \frac{2.34\,g\,N}{0.337\,g\,H} = 6.94$

for compound X, $\frac{g\,N}{g\,H} = \frac{10.96\,g\,N}{1.575\,g\,H} = 6.96$; X is N_2H_4

2.168 $\frac{12.0000}{15.9994} = \frac{X}{16.0000}$; X = 12.0005 for ^{12}C prior to 1961.

2.169 $\frac{39.9626}{15.9994} = \frac{X}{16.0000}$; X = 39.9641 for ^{40}Ca prior to 1961.

2.170 molecular mass = (8 x 12.011) + (9 x 1.0079) + (1 x 14.0067) + (2 x 15.9994) = 151.165 g/mol

2.171 mass % C = $\frac{8 \times 12.011}{151.165} \times 100 = 63.565\%$

mass % H = $\frac{9 \times 1.0079}{151.165} \times 100 = 6.0008\%$

mass % N = $\frac{14.0067}{151.165} \times 100 = 9.2658\%$

mass % O = $\frac{2 \times 15.9994}{151.165} \times 100 = 21.168\%$

2.172 (a) Arrange the droplet charges in order of increasing charge.
2.21×10^{-16} C
4.42×10^{-16} C
4.98×10^{-16} C
6.64×10^{-16} C
7.74×10^{-16} C

Find the smallest charge difference between droplet charges.
$(4.42 - 2.21) \times 10^{-16}$ C = 2.21×10^{-16} C
$(4.98 - 4.42) \times 10^{-16}$ C = 0.56×10^{-16} C
$(6.64 - 4.98) \times 10^{-16}$ C = 1.66×10^{-16} C
$(7.74 - 6.64) \times 10^{-16}$ C = 1.10×10^{-16} C
The smallest difference (0.56×10^{-16} C) is the approximate value for the charge on one blorvek.

To get a more accurate value, divide the droplet charges by the 0.56×10^{-16} C to determine the total number of blorveks.
2.21×10^{-16} C/0.56×10^{-16} C = 4

4.42×10^{-16} C/0.56×10^{-16} C = 8
4.98×10^{-16} C/0.56×10^{-16} C = 9
6.64×10^{-16} C/0.56×10^{-16} C = 12
7.74×10^{-16} C/0.56×10^{-16} C = 14
There are (4 + 8 + 9 + 12 + 14) = 47 blorveks on the 5 droplets with a total charge of $(2.21 + 4.42 + 4.98 + 6.64 + 7.74) \times 10^{-16}$ C = 25.99×10^{-16} C.

The charge on one blorvek is $\frac{25.99 \times 10^{-16}\ \text{C}}{47\ \text{blorveks}} = 5.53 \times 10^{-17}$ C/blorvek.

(b) Again arrange the droplet charges in increasing order.
2.21×10^{-16} C
4.42×10^{-16} C
4.98×10^{-16} C
5.81×10^{-16} C
6.64×10^{-16} C
7.74×10^{-16} C

Find the smallest charge difference between droplet charges.
$(4.42 - 2.21) \times 10^{-16}$ C = 2.21×10^{-16} C
$(4.98 - 4.42) \times 10^{-16}$ C = 0.56×10^{-16} C
$(5.81 - 4.98) \times 10^{-16}$ C = 0.83×10^{-16} C
$(6.64 - 5.81) \times 10^{-16}$ C = 0.83×10^{-16} C
$(7.74 - 6.64) \times 10^{-16}$ C = 1.10×10^{-16} C
The new charge difference (0.83×10^{-16} C) is not an integer multiple of 5.53×10^{-17} C, which means the charge on the blorvek must be smaller than 5.53×10^{-17} C.
The difference between 0.83×10^{-16} C and 0.56×10^{-16} C = 0.27×10^{-16} C is the approximate value for the charge on one blorvek.
To get a more accurate value, divide the droplet charges by the 0.27×10^{-16} C to determine the total number of blorveks.
2.21×10^{-16} C/0.27×10^{-16} C = 8
4.42×10^{-16} C/0.27×10^{-16} C = 16
4.98×10^{-16} C/0.27×10^{-16} C = 18
5.81×10^{-16} C/0.27×10^{-16} C = 21
6.64×10^{-16} C/0.27×10^{-16} C = 24
7.74×10^{-16} C/0.27×10^{-16} C = 28
There are (8 + 16 + 18 + 21 + 24 + 28) = 115 blorveks on the 6 droplets with a total charge of $(2.21 + 4.42 + 4.98 + 5.81 + 6.64 + 7.74) \times 10^{-16}$ C = 31.80×10^{-16} C.

The charge on one blorvek is $\frac{31.80 \times 10^{-16}\ \text{C}}{115\ \text{blorveks}} = 2.77 \times 10^{-17}$ C/blorvek.