2.1 (a) Alpha (α) particles are helium ions with a charge of +2. (b) Beta (β) particles are electrons emitted during the decay of certain radioactive substances. (c) Gamma (γ) rays are high-energy radiation. (d) X-rays are also high-energy radiation, but with a lower energy than γ rays.

2.2 The most common types of radiation known to be emitted by radioactive elements are alpha (α) radiation, beta (β) radiation, and gamma (γ) radiation.

2.3 Alpha (α) particles are composed of two protons and two neutrons. Cathode rays are a stream of electrons. Protons, neutrons, and electrons are fundamental particles. Fundamental particles are particles that were once thought to be the indivisible components of all matter.

2.4 Please see Section 2.2 of the text where early experiments on atomic structure are discussed in detail.

2.5 The sample is emitting particles from its nucleus (α, β, etc.).

2.6 Rutherford used α particles to probe the structure of the atom. Most α particles aimed at thin foils of gold passed through the foil with little or no deflection. A few α particles were deflected at large angles and occasionally an α particle bounced back in the direction from which it had come. Rutherford concluded that most of the atom was empty space with a small, dense, positively charged core (the nucleus).

2.7 First, convert 1 cm to picometers.

\[
1 \text{ cm} \times \frac{0.01 \text{ m}}{1 \text{ cm}} \times \frac{1 \text{ pm}}{1 \times 10^{-12} \text{ m}} = 1 \times 10^{10} \text{ pm}
\]

? He atoms = \( (1 \times 10^{10} \text{ pm}) \times \frac{1 \text{ He atom}}{1 \times 10^2 \text{ pm}} = 1 \times 10^8 \text{ He atoms} \)

2.8 Note that you are given information to set up the unit factor relating meters and miles.

\[
r_{\text{atom}} = 10^4 r_{\text{nucleus}} = 10^4 \times 10 \text{ cm} \times \frac{1 \text{ m}}{100 \text{ cm}} \times \frac{1 \text{ mi}}{1609 \text{ m}} = 0.62 \text{ mi}
\]

2.9 (a) The atomic number is the number of protons in a nucleus. (b) The mass number is the sum of the number of protons and neutrons in the nucleus of an atom. In an atom, the numbers of protons and electrons are equal. Therefore, if the atomic number of an atom is known, both the number of protons and electrons are also known.

2.10 The chemical identity of an atom is determined by its number of protons (atomic number). Isotopes of an element contain differing numbers of neutrons, hence the mass numbers of isotopes of an element will differ. Z is the atomic number, A is the mass number, and X represents the symbol of the element.

2.11 For iron, the atomic number Z is 26. Therefore the mass number A is:

\[
A = 26 + 28 = 54
\]
2.12 **Strategy:** The 239 in Pu-239 is the mass number. The **mass number** \( A \) is the total number of neutrons and protons present in the nucleus of an atom of an element. You can look up the atomic number (number of protons) on the periodic table.

**Solution:**

\[
\text{mass number} = \text{number of protons} + \text{number of neutrons} \\
\text{number of neutrons} = \text{mass number} - \text{number of protons} = 239 - 94 = 145
\]

2.13 | Isotope | \( ^{3/2}\text{He} \) | \( ^{4/2}\text{He} \) | \( ^{24/12}\text{Mg} \) | \( ^{25/12}\text{Mg} \) | \( ^{48/22}\text{Ti} \) | \( ^{79/35}\text{Br} \) | \( ^{195/78}\text{Pt} \) |
<table>
<thead>
<tr>
<th></th>
<th></th>
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<th></th>
<th></th>
<th></th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>No. Protons</td>
<td>2</td>
<td>2</td>
<td>12</td>
<td>12</td>
<td>22</td>
<td>35</td>
<td>78</td>
</tr>
<tr>
<td>No. Neutrons</td>
<td>1</td>
<td>2</td>
<td>12</td>
<td>13</td>
<td>26</td>
<td>44</td>
<td>117</td>
</tr>
</tbody>
</table>

2.14 | Isotope | \( ^{15/7}\text{N} \) | \( ^{33/16}\text{S} \) | \( ^{63/29}\text{Cu} \) | \( ^{84/38}\text{Sr} \) | \( ^{130/56}\text{Ba} \) | \( ^{186/74}\text{W} \) | \( ^{202/80}\text{Hg} \) |
<table>
<thead>
<tr>
<th></th>
<th></th>
<th></th>
<th></th>
<th></th>
<th></th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>No. Protons</td>
<td>7</td>
<td>16</td>
<td>29</td>
<td>38</td>
<td>56</td>
<td>74</td>
<td>80</td>
</tr>
<tr>
<td>No. Neutrons</td>
<td>8</td>
<td>17</td>
<td>34</td>
<td>46</td>
<td>74</td>
<td>112</td>
<td>122</td>
</tr>
<tr>
<td>No. Electrons</td>
<td>7</td>
<td>16</td>
<td>29</td>
<td>38</td>
<td>56</td>
<td>74</td>
<td>80</td>
</tr>
</tbody>
</table>

2.15 (a) \( ^{23/11}\text{Na} \) (b) \( ^{64/28}\text{Ni} \)

2.16 The accepted way to denote the atomic number and mass number of an element \( X \) is as follows:

\[
A^Z_X
\]

where,

\[
A = \text{mass number} \\
Z = \text{atomic number}
\]

(a) \( ^{186/74}\text{W} \) (b) \( ^{201/80}\text{Hg} \)

2.17 Elements can be grouped together according to their chemical and physical properties in a chart called the periodic table. The periodic table enables us to classify elements (as metals, metalloids, and nonmetals) and correlate their properties in a systematic way. Groups are the vertical columns of the periodic table, and periods are the horizontal rows of the table.

2.18 Metals are good conductors of heat and electricity, while nonmetals are usually poor conductors of heat and electricity. Metals, excluding mercury, are solids, whereas many nonmetals are gases.

2.19 (a) Hydrogen (\( \text{H}_2 \)), carbon (\( \text{C} \)), oxygen (\( \text{O}_2 \)), argon (\( \text{Ar} \)).
(b) Sodium (\( \text{Na} \)), titanium (\( \text{Ti} \)), tungsten (\( \text{W} \)), lead (\( \text{Pb} \)).
(c) Silicon (\( \text{Si} \)), germanium (\( \text{Ge} \)), arsenic (\( \text{As} \)), astatine (\( \text{At} \)).

2.20 Column A is the **alkali metals**. Two examples are sodium (\( \text{Na} \)) and potassium (\( \text{K} \)).
Column B is the **alkaline earth metals**. Two examples are calcium (\( \text{Ca} \)) and barium (\( \text{Ba} \)).
Column C is the **halogens**. Two examples are fluorine (\( \text{F} \)) and iodine (\( \text{I} \)).
Column D is the **noble gases**. Two examples are argon (\( \text{Ar} \)) and xenon (\( \text{Xe} \)).

2.21 Helium and Selenium are nonmetals whose name ends with *ium*. (Tellurium is a metalloid whose name ends in *ium.*)
2.22  (a) Metallic character increases as you progress down a group of the periodic table. For example, moving down Group 4A, the nonmetal carbon is at the top and the metal lead is at the bottom of the group.
    (b) Metallic character decreases from the left side of the table (where the metals are located) to the right side of the table (where the nonmetals are located).

2.23  The following data were measured at 20°C.
    (a) Li (0.53 g/cm³)  K (0.86 g/cm³)  H₂O (0.98 g/cm³)
    (b) Au (19.3 g/cm³)  Pt (21.4 g/cm³)  Hg (13.6 g/cm³)
    (c) Os (22.6 g/cm³)
    (d) Te (6.24 g/cm³)

2.24  F and Cl are Group 7A elements; they should have similar chemical properties. Na and K are both Group 1A elements; they should have similar chemical properties. P and N are both Group 5A elements; they should have similar chemical properties.

2.25  An atom is the basic unit of an element that can enter into chemical combination. A molecule is an aggregate of at least two atoms in a definite arrangement held together by chemical forces (also called chemical bonds).

2.26  Allotropes are two or more forms of the same element that differ significantly in chemical and physical properties. Diamond and graphite are allotropes of carbon. Allotropes of an element differ in structure and properties, whereas isotopes of a given element contain different numbers of neutrons but have similar chemistries.

2.27  Two commonly used molecular models are the ball-and-stick model and the space-filling model.

2.28  (a) Na⁺  (b) I⁻  (c) NH₄⁺  (d) SO₄²⁻

2.29  (a) This is a polyatomic molecule that is an elemental form of the substance. It is not a compound.
    (b) This is a polyatomic molecule that is a compound.
    (c) This is a diatomic molecule that is a compound.

2.30  (a) This is a diatomic molecule that is a compound.
    (b) This is a polyatomic molecule that is a compound.
    (c) This is a polyatomic molecule that is the elemental form of the substance. It is not a compound.

2.31  Elements:  N₂, S₈, H₂
    Compounds:  NH₃, NO, CO, CO₂, SO₂

2.32  There are more than two correct answers for each part of the problem.
    (a) H₂ and F₂
    (b) HCl and CO
    (c) S₈ and P₄
    (d) H₂O and C₁₂H₂₂O₁₁ (sucrose)

2.33  Ion  Na⁺  Ca²⁺  Al³⁺  Fe²⁺  I⁻  F⁻  S²⁻  O²⁻  N³⁻
    No. protons 11  20  13  26  53  9  16  8  7
    No. electrons 10  18  10  24  54  10  18  10  10
2.34 The atomic number ($Z$) is the number of protons in the nucleus of each atom of an element. You can find this on a periodic table. The number of electrons in an ion is equal to the number of protons minus the charge on the ion.

\[
\text{number of electrons (ion)} = \text{number of protons} - \text{charge on the ion}
\]

<table>
<thead>
<tr>
<th>Ion</th>
<th>$K^+$</th>
<th>$Mg^{2+}$</th>
<th>$Fe^{3+}$</th>
<th>$Br^-$</th>
<th>$Mn^{2+}$</th>
<th>$C^{4-}$</th>
<th>$Cu^{2+}$</th>
</tr>
</thead>
<tbody>
<tr>
<td>No. protons</td>
<td>19</td>
<td>12</td>
<td>26</td>
<td>35</td>
<td>25</td>
<td>6</td>
<td>29</td>
</tr>
<tr>
<td>No. electrons</td>
<td>18</td>
<td>10</td>
<td>23</td>
<td>36</td>
<td>23</td>
<td>10</td>
<td>27</td>
</tr>
</tbody>
</table>

2.35 Chemical formulas express the composition of molecules and ionic compounds in terms of chemical symbols. (a) 1:1. (b) 1:3. (c) 1:2. (d) 2:3.

2.36 A molecular formula shows the exact number of atoms of each element in the smallest unit of a substance. An empirical formula shows the elements present and the simplest whole-number ratio of the atoms, but not necessarily the actual number of atoms in a given molecule.

2.37 Cyclobutane ($C_4H_8$) and cyclohexane ($C_6H_{12}$) have the same empirical formula ($CH_2$), but different molecular formulas.

2.38 $P_4$ signifies one molecule that is composed of four P atoms. 4P represents four atoms of P (phosphorus).

2.39 An ionic compound contains cations and anions. Electrical neutrality is maintained because the positive charge of the cations is balanced by the negative charge of the anions.

2.40 Ionic compounds do not consist of discrete molecular units, but are three-dimensional networks of ions. The formula of ionic compounds represents the simplest ratio (empirical formula) in which the cation and anion combine.

2.41 (a) CN  (b) CH  (c) C$_9$H$_{20}$  (d) P$_2$O$_5$  (e) BH$_3$

2.42 **Strategy:** An empirical formula tells us which elements are present and the simplest whole-number ratio of their atoms. Can you divide the subscripts in the formula by some factor to end up with smaller whole-number subscripts?

**Solution:**

(a) Dividing both subscripts by 2, the simplest whole number ratio of the atoms in Al$_2$Br$_6$ is AlBr$_3$.

(b) Dividing all subscripts by 2, the simplest whole number ratio of the atoms in Na$_2$S$_2$O$_4$ is NaSO$_2$.

(c) The molecular formula as written, N$_2$O$_4$, contains the simplest whole number ratio of the atoms present. In this case, the molecular formula and the empirical formula are the same.

(d) The molecular formula as written, K$_2$Cr$_2$O$_7$, contains the simplest whole number ratio of the atoms present. In this case, the molecular formula and the empirical formula are the same.

2.43 The molecular formula of glycine is C$_2$H$_5$NO$_2$.

2.44 The molecular formula of ethanol is C$_2$H$_6$O.
2.45 Compounds of metals with nonmetals are usually ionic. Nonmetal-nonmetal compounds are usually molecular.

Ionic: \( \text{LiF, BaCl}_2, \text{KCl} \)
Molecular: \( \text{SiCl}_4, \text{B}_2\text{H}_6, \text{C}_2\text{H}_4 \)

2.46 Compounds of metals with nonmetals are usually ionic. Nonmetal-nonmetal compounds are usually molecular.

Ionic: \( \text{NaBr, BaF}_2, \text{CsCl} \)
Molecular: \( \text{CH}_4, \text{CCl}_4, \text{ICl}, \text{NF}_3 \)

2.47 (a) sodium chromate  
(b) potassium hydrogen phosphate  
(c) hydrogen bromide (molecular compound)  
(d) hydrobromic acid  
(e) lithium carbonate  
(f) potassium dichromate  
(g) ammonium nitrite  
(h) phosphorus trifluoride  
(i) phosphorus pentafluoride  
(j) tetraphosphorus hexoxide  
(k) cadmium iodide  
(l) strontium sulfate  
(m) aluminum hydroxide  
(n) sodium carbonate decahydrate

2.48 Strategy: When naming ionic compounds, our reference for the names of cations and anions is Table 2.3 of the text. Keep in mind that if a metal can form cations of different charges, we need to use the Stock system. In the Stock system, Roman numerals are used to specify the charge of the cation. The metals that have only one charge in ionic compounds are the alkali metals \(+1\), the alkaline earth metals \(+2\), \(\text{Ag}^+, \text{Zn}^{2+}, \text{Cd}^{2+}, \text{and Al}^{3+}\).

When naming acids, binary acids are named differently than oxoacids. For binary acids, the name is based on the nonmetal. For oxoacids, the name is based on the polyatomic anion. For more detail, see Section 2.7 of the text.

Solution:

(a) This is an ionic compound in which the metal cation \(\text{K}^+\) has only one charge. The correct name is potassium hypochlorite. Hypochlorite is a polyatomic ion with one less \(\text{O}\) atom than the chlorite ion, \(\text{ClO}_2^-\)

(b) silver carbonate

(c) This is an ionic compound in which the metal can form more than one cation. Use a Roman numeral to specify the charge of the \(\text{Fe}^+\) ion. Since each chloride ion has a \(-1\) charge, the \(\text{Fe}^+\) ion has a \(+2\) charge. The correct name is iron(II) chloride.

(d) potassium permanganate  
(e) cesium chloride  
(f) hypoiodous acid  
(g) This is an ionic compound in which the metal can form more than one cation. Use a Roman numeral to specify the charge of the \(\text{Fe}^+\) ion. Since the oxide ion has a \(-2\) charge, the \(\text{Fe}^+\) ion has a \(+2\) charge. The correct name is iron(II) oxide.

(h) iron(III) oxide  
(i) This is an ionic compound in which the metal can form more than one cation. Use a Roman numeral to specify the charge of the \(\text{Ti}^+\) ion. Since each of the four chloride ions has a \(-1\) charge (total of \(-4\)), the \(\text{Ti}^+\) ion has a \(+4\) charge. The correct name is titanium(IV) chloride.

(j) sodium hydride  
(k) lithium nitride  
(l) sodium oxide
(m) This is an ionic compound in which the metal cation (Na\(^+\)) has only one charge. The O\(_2\)\(^{2-}\) ion is called the peroxide ion. Each oxygen has a \(-1\) charge. You can determine that each oxygen only has a \(-1\) charge, because each of the two Na ions has a \(+1\) charge. Compare this to sodium oxide in part (l). The correct name is sodium peroxide.

(n) This is an ionic compound in which the metal can form more than one cation. Use a Roman numeral to specify the charge of the Fe ion. Since each chloride ion has a \(-1\) charge, the Fe ion has a \(+3\) charge. At the end of the name, we add hexahydrate for the six waters of hydration. The correct name is iron(III) chloride hexahydrate.

2.49  
(a) RbNO\(_2\)  
(b) K\(_2\)S  
(c) HBrO\(_4\)  
(d) Mg\(_3\)(PO\(_4\))\(_2\)  
(e) CaHPO\(_4\)  
(f) BCl\(_3\)  
(g) IF\(_7\)  
(h) (NH\(_4\))\(_2\)SO\(_4\)  
(i) AgClO\(_4\)  
(j) Fe\(_2\)(CrO\(_4\))\(_3\)  
(k) CaSO\(_4\)\(\cdot\)2H\(_2\)O

2.50 **Strategy:** When writing formulas of molecular compounds, the prefixes specify the number of each type of atom in the compound.

When writing formulas of ionic compounds, the subscript of the cation is numerically equal to the charge of the anion, and the subscript of the anion is numerically equal to the charge on the cation. If the charges of the cation and anion are numerically equal, then no subscripts are necessary. Charges of common cations and anions are listed in Table 2.3 of the text. Keep in mind that Roman numerals specify the charge of the cation, not the number of metal atoms. Remember that a Roman numeral is not needed for some metal cations, because the charge is known. These metals are the alkali metals (+1), the alkaline earth metals (+2), Ag\(^+\), Zn\(^{2+}\), Cd\(^{2+}\), and Al\(^{3+}\).

When writing formulas of oxoacids, you must know the names and formulas of polyatomic anions (see Table 2.3 of the text).

**Solution:**

(a) The Roman numeral I tells you that the Cu cation has a \(+1\) charge. Cyanide has a \(-1\) charge. Since, the charges are numerically equal, no subscripts are necessary in the formula. The correct formula is CuCN.

(b) Strontium is an alkaline earth metal. It only forms a \(+2\) cation. The polyatomic ion chlorite, ClO\(_2\)\(^-\), has a \(-1\) charge. Since the charges on the cation and anion are numerically different, the subscript of the cation is numerically equal to the charge on the anion, and the subscript of the anion is numerically equal to the charge on the cation. The correct formula is Sr(ClO\(_2\))\(_2\).

(c) Perchloric tells you that the anion of this oxoacid is perchlorate, ClO\(_4\)^-. The correct formula is HClO\(_4\)(aq). Remember that (aq) means that the substance is dissolved in water.

(d) Hydroiodic tells you that the anion of this binary acid is iodide, I\(^-\). The correct formula is HI(aq).

(e) Na is an alkali metal. It only forms a \(+1\) cation. The polyatomic ion ammonium, NH\(_4\)^+, has a \(+1\) charge and the polyatomic ion phosphate, PO\(_4\)^{3-}, has a \(-3\) charge. To balance the charge, you need 2 Na\(^+\) cations. The correct formula is Na\(_2\)(NH\(_4\))PO\(_4\).

(f) The Roman numeral II tells you that the Pb cation has a \(+2\) charge. The polyatomic ion carbonate, CO\(_3\)^{2-}, has a \(-2\) charge. Since, the charges are numerically equal, no subscripts are necessary in the formula. The correct formula is PbCO\(_3\).

(g) The Roman numeral II tells you that the Sn cation has a \(+2\) charge. Fluoride has a \(-1\) charge. Since the charges on the cation and anion are numerically different, the subscript of the cation is numerically equal to the charge on the anion, and the subscript of the anion is numerically equal to the charge on the cation. The correct formula is SnF\(_2\).

(h) This is a molecular compound. The Greek prefixes tell you the number of each type of atom in the molecule. The correct formula is P\(_4\)S\(_{10}\).
(i) The Roman numeral II tells you that the Hg cation has a +2 charge. Oxide has a −2 charge. Since, the charges are numerically equal, no subscripts are necessary in the formula. The correct formula is **HgO**.

(j) The Roman numeral I tells you that the Hg cation has a +1 charge. However, this cation exists as Hg\(^{2+}\). Iodide has a −1 charge. You need two iodide ions to balance the +2 charge of Hg\(^{2+}\). The correct formula is **Hg\(_2\)I\(_2\)**.

(k) The Roman numeral II tells you that the Co cation has a +2 charge. Chloride has a −1 charge. Since the charges on the cation and anion are numerically different, the subscript of the cation is numerically equal to the charge on the anion, and the subscript of the anion is numerically equal to the charge on the cation. We add 6H\(_2\)O at the end of the formula for the six waters of hydration. The correct formula is **CoCl\(_2\)\cdot 6H\(_2\)O**.

2.51 The number of protons = 65 − 35 = 30. The element that contains 30 protons is zinc, Zn. There are two fewer electrons than protons, so the charge of the cation is +2. The symbol for this cation is **Zn\(^{2+}\)**.

2.52 Changing the electrical charge of an atom usually has a major effect on its chemical properties. The two electrically neutral carbon isotopes should have nearly identical chemical properties; that is (c).

2.53 (a) Species with the same number of protons and electrons will be neutral. A, F, G.

(b) Species with more electrons than protons will have a negative charge. B, E.

(c) Species with more protons than electrons will have a positive charge. C, D.

(d) A: \(^{10}\text{B}\) B: \(^{14}\text{N}\) C: \(^{39}\text{K}\) D: \(^{66}\text{Zn}\) E: \(^{81}\text{Br}\) F: \(^{11}\text{B}\) G: \(^{19}\text{F}\)

2.54 (a) Does this refer to hydrogen atoms or hydrogen molecules? One can’t be sure.

(b) NaCl is an ionic compound; it doesn’t form molecules.

2.55 Yes. The law of multiple proportions requires that the masses of sulfur combining with phosphorus must be in the ratios of small whole numbers. For the three compounds shown, four phosphorus atoms combine with three, seven, and ten sulfur atoms, respectively. If the atom ratios are in small whole number ratios, then the mass ratios must also be in small whole number ratios.

2.56 The species and their identification are as follows:

(a) **SO\(_2\)** molecule and compound

(b) **S\(_8\)** element and molecule

(c) **Cs** element

(d) **N\(_2\)O\(_5\)** molecule and compound

(e) **O** element

(f) **O\(_2\)** element and molecule

(g) **O\(_3\)** element and molecule

(h) **CH\(_4\)** molecule and compound

(i) **KBr** compound

(j) **S** element

(k) **P\(_4\)** element and molecule

(l) **LiF** compound

2.57 (a) molecular, C\(_3\)H\(_8\) (b) molecular, C\(_2\)H\(_2\) (c) molecular, C\(_2\)H\(_6\) (d) molecular, C\(_6\)H\(_6\) (e) empirical, C\(_3\)H\(_8\) (f) empirical, CH (g) empirical, CH\(_3\)

2.58 (a) CO\(_2\) (s), solid carbon dioxide

(b) NaCl, sodium chloride

(c) N\(_2\)O, nitrous oxide

(d) CaCO\(_3\), calcium carbonate

(e) CaO, calcium oxide

(f) Ca(OH)\(_2\), calcium hydroxide

(g) NaHCO\(_3\), sodium bicarbonate

(h) Mg(OH)\(_2\), magnesium hydroxide
2.59 Symbol  \( ^{11}B \)  \( ^{54}Fe^{2+} \)  \( ^{31}P^{3-} \)  \( ^{196}Au \)  \( ^{222}Rn \)
Protons 5 26 15 79 86
Neutrons 6 28 16 117 136
Electrons 5 24 18 79 86
Net Charge 0 +2 −3 0 0

2.60 (a) Ionic compounds are typically formed between metallic (especially Groups 1A, 2A, and aluminum) and nonmetallic elements.
(b) In general the transition metals, the actinides and lanthanides have variable charges.

2.61 Group 1A metals form \( M^+ \) ions. Group 2A metals form \( Y^{2+} \) ions. Aluminum forms an \( Al^{3+} \) ion. Oxygen forms an \( O^{2-} \) ion (oxide). Nitrogen forms an \( N^{3-} \) ion (nitride), and the halogens form \( X^- \) ions. Making a table:

<table>
<thead>
<tr>
<th>Nonmetals</th>
<th>1A Metals</th>
<th>2A Metals</th>
<th>Aluminum</th>
</tr>
</thead>
<tbody>
<tr>
<td>Halogens</td>
<td>MX</td>
<td>YX₂</td>
<td>AlX₃</td>
</tr>
<tr>
<td>Oxygen</td>
<td>M₂O</td>
<td>YO</td>
<td>Al₂O₃</td>
</tr>
<tr>
<td>Nitrogen</td>
<td>M₃N</td>
<td>Y₃N₂</td>
<td>AlN</td>
</tr>
</tbody>
</table>

2.62 The symbol \( ^{23}Na \) provides more information than \( ^{11}Na \). The mass number plus the chemical symbol identifies a specific isotope of Na (sodium) while combining the atomic number with the chemical symbol tells you nothing new. Can other isotopes of sodium have different atomic numbers?

2.63 The binary Group 7A element acids are: HF, hydrofluoric acid; HCl, hydrochloric acid; HBr, hydrobromic acid; HI, hydroiodic acid. Oxoacids containing Group 7A elements (using the specific examples for chlorine) are: HClO₄, perchloric acid; HClO₃, chloric acid; HClO₂, chlorous acid: HClO, hypochlorous acid.

Examples of oxoacids containing other Group A-block elements are: H₃BO₃, boric acid (Group 3A); H₂CO₃, carbonic acid (Group 4A); HNO₃, nitric acid and H₃PO₄, phosphoric acid (Group 5A); and H₂SO₄, sulfuric acid (Group 6A). Hydrosulfuric acid, H₂S, is an example of a binary Group 6A acid while HCN, hydrocyanic acid, contains both a Group 4A and 5A element.

2.64 Isotope: \( ^{40}Mg \)  \( ^{44}Si \)  \( ^{48}Ca \)  \( ^{43}Al \)
No. Neutrons: 28 30 28 30

2.65 F and Cl are Group 7A elements; they should have similar chemical properties. Na and K are both Group 1A elements; they should have similar chemical properties. P and N are both Group 5A elements; they should have similar chemical properties.

2.66 H₂, N₂, O₂, F₂, Cl₂, He, Ne, Ar, Kr, Xe, Rn

2.67 Cu, Ag, and Au are fairly chemically unreactive. This makes them specially suitable for making coins and jewelry, that you want to last a very long time.

2.68 They do not have a strong tendency to form compounds. Helium and neon are chemically inert.
2.69 Magnesium is alkaline earth metal like calcium. You should expect the charge of the metal to be the same (+2). The formula of magnesium oxide is \( \text{MgO} \). Cesium is an alkali metal. The charge of an alkali metal in an ionic compound is +1. The formula of cesium oxide is \( \text{Cs}_2\text{O} \).

2.70 All isotopes of radium are radioactive. It is a radioactive decay product of uranium-238. Radium itself does not occur naturally on Earth.

2.71 The mass of fluorine reacting with hydrogen and deuterium would be the same. The ratio of F atom to hydrogen (or deuterium) is 1:1 in both compounds. This does not violate the law of definite proportions. When the law of definite proportions was formulated, scientists did not know of the existence of isotopes.

2.72 (a) NaH, sodium hydride  (b) \( \text{B}_2\text{O}_3 \), diboron trioxide  (c) \( \text{Na}_2\text{S} \), sodium sulfide  
   (d) AlF\(_3\), aluminum fluoride  (e) \( \text{OF}_2 \), oxygen difluoride  (f) SrCl\(_2\), strontium chloride

2.73

<table>
<thead>
<tr>
<th>Cation</th>
<th>Anion</th>
<th>Formula</th>
<th>Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>( \text{Mg}^{2+} )</td>
<td>( \text{HCO}_3^- )</td>
<td>( \text{Mg(HCO}_3)_2 )</td>
<td>Magnesium bicarbonate</td>
</tr>
<tr>
<td>( \text{Sr}^{2+} )</td>
<td>( \text{Cl}^- )</td>
<td>( \text{SrCl}_2 )</td>
<td>Strontium chloride</td>
</tr>
<tr>
<td>( \text{Fe}^{3+} )</td>
<td>( \text{NO}_2^- )</td>
<td>( \text{Fe(NO}_2)_3 )</td>
<td>Iron(III) nitrite</td>
</tr>
<tr>
<td>( \text{Mn}^{2+} )</td>
<td>( \text{ClO}_3^- )</td>
<td>( \text{Mn(ClO}_3)_2 )</td>
<td>Manganese(II) chlorate</td>
</tr>
<tr>
<td>( \text{Sn}^{4+} )</td>
<td>( \text{Br}^- )</td>
<td>( \text{SnBr}_4 )</td>
<td>Tin(IV) bromide</td>
</tr>
<tr>
<td>( \text{Co}^{2+} )</td>
<td>( \text{PO}_4^{3-} )</td>
<td>( \text{Co}_3(\text{PO}_4)_2 )</td>
<td>Cobalt(II) phosphate</td>
</tr>
<tr>
<td>( \text{Hg}_2^{2+} )</td>
<td>( \text{I}^- )</td>
<td>( \text{Hg}_2\text{I}_2 )</td>
<td>Mercury(I) iodide</td>
</tr>
<tr>
<td>( \text{Cu}^{2+} )</td>
<td>( \text{CO}_3^{2-} )</td>
<td>( \text{Cu}_2\text{CO}_3 )</td>
<td>Copper(I) carbonate</td>
</tr>
<tr>
<td>( \text{Li}^+ )</td>
<td>( \text{N}^{3-} )</td>
<td>( \text{Li}_3\text{N} )</td>
<td>Lithium nitride</td>
</tr>
<tr>
<td>( \text{Al}^{3+} )</td>
<td>( \text{S}^{2-} )</td>
<td>( \text{Al}_2\text{S}_3 )</td>
<td>Aluminum sulfide</td>
</tr>
</tbody>
</table>

2.74 (a) Br  (b) Rn  (c) Se  (d) Rb  (e) Pb

2.75 \( \text{NF}_3 \), nitrogen trifluoride  
\( \text{PBr}_5 \), phosphorus pentabromide  
\( \text{SCl}_2 \), sulfur dichloride

2.76 The change in energy is equal to the energy released. We call this \( \Delta E \). Similarly, \( \Delta m \) is the change in mass. Because \( m = \frac{E}{c^2} \), we have

\[
\Delta m = \frac{\Delta E}{c^2} = \frac{(1.715 \times 10^3 \text{ kJ}) \times \frac{1000 \text{ J}}{1 \text{ kJ}}}{(3.00 \times 10^8 \text{ m/s})^2} = 1.91 \times 10^{-11} \text{ kg} = 1.91 \times 10^{-8} \text{ g}
\]

Note that we need to convert kJ to J so that we end up with units of kg for the mass. \( 1 \text{ J} = \frac{1 \text{ kg} \cdot \text{m}^2}{s^2} \)

We can add together the masses of hydrogen and oxygen to calculate the mass of water that should be formed.

\[
12.096 \text{ g} + 96.000 \text{ g} = 108.096 \text{ g}
\]
The predicted change (loss) in mass is only $1.91 \times 10^{-8}$ g which is too small a quantity to measure accurately. Therefore, for all practical purposes, the law of conservation of mass is assumed to hold for ordinary chemical processes.

2.77 (a) Rutherford’s experiment is described in detail in Section 2.2 of the text. From the average magnitude of scattering, Rutherford estimated the number of protons (based on electrostatic interactions) in the nucleus.

(b) Assuming that the nucleus is spherical, the volume of the nucleus is:

$$ V = \frac{4}{3} \pi r^3 = \frac{4}{3} \pi (3.04 \times 10^{-13} \text{ cm})^3 = 1.18 \times 10^{-37} \text{ cm}^3 $$

The density of the nucleus can now be calculated.

$$ d = \frac{m}{V} = \frac{3.82 \times 10^{-23} \text{ g}}{1.18 \times 10^{-37} \text{ cm}^3} = 3.24 \times 10^{14} \text{ g/cm}^3 $$

To calculate the density of the space occupied by the electrons, we need both the mass of 11 electrons, and the volume occupied by these electrons.

The mass of 11 electrons is:

$$ 11 \text{ electrons} \times \frac{9.1095 \times 10^{-28} \text{ g}}{1 \text{ electron}} = 1.0020 \times 10^{-26} \text{ g} $$

The volume occupied by the electrons will be the difference between the volume of the atom and the volume of the nucleus. The volume of the nucleus was calculated above. The volume of the atom is calculated as follows:

$$ 186 \text{ pm} \times \frac{1 \times 10^{-12} \text{ m}}{1 \text{ pm}} \times \frac{1 \text{ cm}}{1 \times 10^{-2} \text{ m}} = 1.86 \times 10^{-8} \text{ cm} $$

$$ V_{\text{atom}} = \frac{4}{3} \pi r^3 = \frac{4}{3} \pi (1.86 \times 10^{-8} \text{ cm})^3 = 2.70 \times 10^{-23} \text{ cm}^3 $$

$$ V_{\text{electrons}} = V_{\text{atom}} - V_{\text{nucleus}} = (2.70 \times 10^{-23} \text{ cm}^3) - (1.18 \times 10^{-37} \text{ cm}^3) = 2.70 \times 10^{-23} \text{ cm}^3 $$

As you can see, the volume occupied by the nucleus is insignificant compared to the space occupied by the electrons.

The density of the space occupied by the electrons can now be calculated.

$$ d = \frac{m}{V} = \frac{1.0020 \times 10^{-26} \text{ g}}{2.70 \times 10^{-23} \text{ cm}^3} = 3.71 \times 10^{-4} \text{ g/cm}^3 $$

The above results do support Rutherford's model. Comparing the space occupied by the electrons to the volume of the nucleus, it is clear that most of the atom is empty space. Rutherford also proposed that the nucleus was a dense central core with most of the mass of the atom concentrated in it. Comparing the density of the nucleus with the density of the space occupied by the electrons also supports Rutherford's model.
2.78 (a) Ethane Acetylene
2.65 g C 4.56 g C
0.665 g H 0.383 g H

Let’s compare the ratio of the hydrogen masses in the two compounds. To do this, we need to start with the same mass of carbon. If we were to start with 4.56 g of C in ethane, how much hydrogen would combine with 4.56 g of carbon?

\[ 0.665 \text{ g H} \times \frac{4.56 \text{ g C}}{2.65 \text{ g C}} = 1.14 \text{ g H} \]

We can calculate the ratio of H in the two compounds.

\[ \frac{1.14 \text{ g}}{0.383 \text{ g}} \approx 3 \]

This is consistent with the Law of Multiple Proportions which states that if two elements combine to form more than one compound, the masses of one element that combine with a fixed mass of the other element are in ratios of small whole numbers. In this case, the ratio of the masses of hydrogen in the two compounds is 3:1.

(b) For a given amount of carbon, there is 3 times the amount of hydrogen in ethane compared to acetylene. Reasonable formulas would be:

<table>
<thead>
<tr>
<th></th>
<th>Ethane</th>
<th>Acetylene</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>CH₃</td>
<td>CHC₂H₂</td>
</tr>
<tr>
<td></td>
<td>C₂H₆</td>
<td>C₂H₂</td>
</tr>
</tbody>
</table>

2.79 Two different structural formulas for the molecular formula C₂H₆O are:

\[ \text{H–C–C–O–H} \quad \text{H–C–O–C–H} \]

In the second hypothesis of Dalton’s Atomic Theory, he states that in any compound, the ratio of the number of atoms of any two of the elements present is either an integer or simple fraction. In the above two compounds, the ratio of atoms is the same. This does not necessarily contradict Dalton’s hypothesis, but Dalton was not aware of chemical bond formation and structural formulas.

2.80 The mass number is the sum of the number of protons and neutrons in the nucleus.

\[ \text{Mass number} = \text{number of protons} + \text{number of neutrons} \]

Let the atomic number (number of protons) equal \( A \). The number of neutrons will be \( 1.2A \). Plug into the above equation and solve for \( A \).

\[ 55 = A + 1.2A \]
\[ A = 25 \]

The element with atomic number 25 is manganese, Mn.

2.81 The acids, from left to right, are chloric acid, nitrous acid, hydrocyanic acid, and sulfuric acid.