CHAPTER 2 THE COMPONENTS OF MATTER

TOOL OF THE LABORATORY BOXED READING PROBLEMS

B2.1 Plan: There is one peak for each type of Cl atom and peaks for the Cl_2 molecule. The *m*/*e* ratio equals the mass divided by 1+.
Solution:
a) There is one peak for the ^{35}Cl atom and another peak for the ^{37}Cl atom. There are three peaks for the three possible Cl_2 molecules: $^{35}Cl^{35}Cl$ (both atoms are mass 35), $^{37}Cl^{37}Cl$ (both atoms are mass 37), and $^{35}Cl^{37}Cl$ (one atom is mass 35 and one is mass 37). So the mass of chlorine will have **5 peaks**.
b)

Peak	*m*/*e* ratio	
^{35}Cl	**35**	**lightest particle**
^{37}Cl	37	
$^{35}Cl^{35}Cl$	70 (35 + 35)	
$^{35}Cl^{37}Cl$	72 (35 + 37)	
$^{37}Cl^{37}Cl$	**74** (35 + 37)	**heaviest particle**

B2.2 Plan: Each peak in the mass spectrum of carbon represents a different isotope of carbon. The heights of the peaks correspond to the natural abundances of the isotopes.
Solution:
Carbon has three naturally occurring isotopes: ^{12}C, ^{13}C, and ^{14}C. ^{12}C has an abundance of 98.89% and would have the tallest peak in the mass spectrum as the most abundant isotope. ^{13}C has an abundance of 1.11% and thus would have a significantly shorter peak; the shortest peak in the mass spectrum would correspond to the least abundant isotope, ^{14}C, the abundance of which is less than 0.01%. Peak Y, as the tallest peak, has a *m*/*e* ratio of 12 (^{12}C); X, the shortest peak, has a *m*/*e* ratio of 14(^{14}C). Peak Z corresponds to ^{13}C with a *m*/*e* ratio of **13**.

B2.3 Plan: Review the discussion on separations.
Solution:
a) Salt dissolves in water and pepper does not. Procedure: add water to mixture and filter to remove solid pepper. Evaporate water to recover solid salt.
b) The water/soot mixture can be filtered; the water will flow through the filter paper, leaving the soot collected on the filter paper.
c) Allow the mixture to warm up, and then pour off the melted ice (water); or, add water, and the glass will sink and the ice will float.
d) Heat the mixture; the alcohol will boil off (distill), while the sugar will remain behind.
e) The spinach leaves can be extracted with a solvent that dissolves the pigments. Chromatography can be used to separate one pigment from the other.

END–OF–CHAPTER PROBLEMS

2.1 Plan: Refer to the definitions of an element and a compound.
Solution:
Unlike compounds, elements cannot be broken down by chemical changes into simpler materials. Compounds contain different types of atoms; there is only one type of atom in an element.

2.2 Plan: Refer to the definitions of a compound and a mixture.
Solution:
1) A compound has constant composition but a mixture has variable composition. 2) A compound has distinctly different properties than its component elements; the components in a mixture retain their individual properties.

2.3 Plan: Recall that a substance has a fixed composition.
Solution:
a) The fixed mass ratio means it has constant composition, thus, it is a **pure substance** (compound).
b) All the atoms are identical, thus, it is a **pure substance** (element).
c) The composition can vary, thus, this is an **impure substance** (a mixture).
d) The specific arrangement of different atoms means it has constant composition, thus, it is a **pure substance** (compound).

2.4 Plan: Remember that an element contains only one kind of atom while a compound contains at least two different elements (two kinds of atoms) in a fixed ratio. A mixture contains at least two different substances in a composition that can vary.
Solution:
a) The presence of more than one element (calcium and chlorine) makes this pure substance a **compound**.
b) There are only atoms from one element, sulfur, so this pure substance is an **element**.
c) This is a combination of two compounds and has a varying composition, so this is a **mixture**.
d) The presence of more than one type of atom means it cannot be an element. The specific, not variable, arrangement means it is a **compound**.

2.5 Some elements, such as the noble gases (He, Ne, Ar, etc.) occur as individual atoms. Many other elements, such as most other nonmetals (O_2, N_2, S_8, P_4, etc.) occur as molecules.

2.6 Compounds contain atoms from two or more elements, thus the smallest unit must contain at least a pair of atoms in a molecule.

2.7 Mixtures have variable composition; therefore, the amounts may vary. Compounds, as pure substances, have constant composition so their composition cannot vary.

2.8 The tap water must be a mixture, since it consists of some unknown (and almost certainly variable) amount of dissolved substance in solution in the water.

2.9 Plan: Recall that an element contains only one kind of atom; the atoms in an element may occur as molecules. A compound contains two kinds of atoms (different elements).
Solution:
a) This scene has 3 atoms of an element, 2 molecules of one compound (with one atom each of two different elements), and 2 molecules of a second compound (with 2 atoms of one element and one atom of a second element).
b) This scene has 2 atoms of one element, 2 molecules of a diatomic element, and 2 molecules of a compound (with one atom each of two different elements).
c) This scene has 2 molecules composed of 3 atoms of one element and 3 diatomic molecules of the same element.

2.10 Plan: Recall that a mixture is composed of two or more substances physically mixed, with a composition that can vary.
Solution:
The street sample is a mixture. The mass of vitamin C per gram of drug sample can vary. Therefore, if several samples of the drug have the same mass of vitamin C per gram of sample, this is an indication that the samples all have a common source. Samples of the street drugs with varying amounts of vitamin C per gram of sample have different sources. The constant mass ratio of the components indicates mixtures that have the same composition by accident, not of necessity.

2.11 Separation techniques allow mixtures (with varying composition) to be separated into the pure substance components which can then be analyzed by some method. Only when there is a reliable way of determining the composition of a sample, can you determine if the composition is constant.

2.12 Plan: Restate the three laws in your own words.
Solution:
a) The law of mass conservation applies to all substances — **elements, compounds, and mixtures**. Matter can neither be created nor destroyed, whether it is an element, compound, or mixture.
b) The law of definite composition applies to **compounds** only, because it refers to a constant, or definite, composition of elements within a compound.
c) The law of multiple proportions applies to **compounds** only, because it refers to the combination of elements to form compounds.

2.13 In ordinary chemical reactions (i.e., those that do not involve nuclear transformations), mass is conserved and the law of mass conservation is still valid.

2.14 Plan: Review the three laws: law of mass conservation, law of definite composition, and law of multiple proportions.
Solution:
a) **Law of Definite Composition** — The compound potassium chloride, KCl, is composed of the same elements and same fraction by mass, regardless of its source (Chile or Poland).
b) **Law of Mass Conservation** — The mass of the substances inside the glass bulb did not change during the chemical reaction (formation of magnesium oxide from magnesium and oxygen).
c) **Law of Multiple Proportions —** Two elements, O and As, can combine to form two different compounds that have different proportions of As present.

2.15 Plan: The law of multiple proportions states that two elements can form two different compounds in which the proportions of the elements are different.
Solution:
Scene B illustrates the law of multiple proportions for compounds of chlorine and oxygen. The law of multiple proportions refers to the different compounds that two elements can form that have different proportions of the elements. Scene B shows that chlorine and oxygen can form both Cl_2O, dichlorine monoxide, and ClO_2, chlorine dioxide.

2.16 Plan: Review the definition of percent by mass.
Solution:
a) **No**, the mass percent of each element in a compound is fixed. The percentage of Na in the compound NaCl is 39.34% (22.99 u/58.44 u), whether the sample is 0.5000 g or 50.00 g.
b) **Yes**, the mass of each element in a compound depends on the mass of the compound. A 0.5000 g sample of NaCl contains 0.1967 g of Na (39.34% of 0.5000 g), whereas a 50.00 g sample of NaCl contains 19.67 g of Na (39.34% of 50.00 g).

2.17 Generally no, the composition of a compound is determined by the elements used, not their amounts. If too much of one element is used, the excess will remain as unreacted element when the reaction is over.

2.18 Plan: Review the mass laws: law of mass conservation, law of definite composition, and law of multiple proportions. For each experiment, compare the mass values before and after each reaction and examine the ratios of the mass of white compound to the mass of colourless gas.
Solution:
Experiment 1: mass before reaction = 1.00 g; mass after reaction = 0.64 g + 0.36 g = 1.00 g
Experiment 2: mass before reaction = 3.25 g; mass after reaction = 2.08 g + 1.17 g = 3.25 g
Both experiments demonstrate the **law of mass conservation** since the total mass before reaction equals the total mass after reaction.
Experiment 1: mass white compound/mass colourless gas = 0.64 g/0.36 g = 1.78
Experiment 2: mass white compound/mass coluorless gas = 2.08 g/1.17 g = 1.78
Both Experiments 1 and 2 demonstrate the **law of definite composition** since the compound has the same composition by mass in each experiment.

2.19 Plan: Review the mass laws: law of mass conservation, law of definite composition, and law of multiple proportions. For each experiment, compare the mass values before and after each reaction and examine the ratios of the mass of reacted copper to the mass of reacted iodine.

Solution:

Experiment 1: mass before reaction = 1.27 g + 3.50 g = 4.77 g; mass after reaction = 3.81 g + 0.96 g = 4.77 g
Experiment 2: mass before reaction = 2.55 g + 3.50 g = 6.05 g; mass after reaction = 5.25 g + 0.80 g = 6.05 g
Both experiments demonstrate the **law of mass conversation** since the total mass before reaction equals the total mass after reaction.

Experiment 1: mass of reacted copper = 1.27 g; mass of reacted iodine = 3.50 g – 0.96 g = 2.54 g
Mass reacted copper/mass reacted iodine = 1.27 g/2.54 g = 0.50
Experiment 2: mass of reacted copper = 2.55 g – 0.80 g = 1.75 g; mass of reacted iodine = 3.50 g
Mass reacted copper/mass reacted iodine = 1.75 g/3.50 g = 0.50
Both Experiments 1 and 2 demonstrate the **law of definite composition** since the compound has the same composition by mass in each experiment.

2.20 Plan: Fluorite is a mineral containing only calcium and fluorine. The difference between the mass of fluorite and the mass of calcium gives the mass of fluorine. Mass fraction is calculated by dividing the mass of element by the mass of compound (fluorite) and mass percent is obtained by multiplying the mass fraction by 100.

Solution:

a) Mass (g) of fluorine = mass of fluorite – mass of calcium = 2.76 g – 1.42 g = **1.34 g fluorine**

b) Mass fraction of Ca = $\frac{\text{mass Ca}}{\text{mass fluorite}} = \frac{1.42 \text{ g Ca}}{2.76 \text{ g fluorite}}$ = 0.51449 = **0.514**

Mass fraction of F = $\frac{\text{mass F}}{\text{mass fluorite}} = \frac{1.34 \text{ g F}}{2.76 \text{ g fluorite}}$ = 0.48551 = **0.486**

c) Mass percent of Ca = 0.51449 x 100 %= 51.449 %= **51.4%**
Mass percent of F = 0.48551 x 100 %= 48.551 %= **48.6%**

2.21 Plan: Galena is a mineral containing only lead and sulfur. The difference between the mass of galena and the mass of lead gives the mass of sulfur. Mass fraction is calculated by dividing the mass of element by the mass of compound (galena) and mass percent is obtained by multiplying the mass fraction by 100.

Solution:

a) Mass (g) of sulfur = mass of galena – mass of sulfur = 2.34 g – 2.03 g = **0.31 g sulfur**

b) Mass fraction of Pb = $\frac{\text{mass Pb}}{\text{mass galena}} = \frac{2.03 \text{ g Pb}}{2.34 \text{ g galena}}$ = 0.8675214 = **0.868**

Mass fraction of S = $\frac{\text{mass S}}{\text{mass galena}} = \frac{0.31 \text{ g S}}{2.34 \text{ g galena}}$ = 0.1324786 = **0.13**

c) Mass percent of Pb = (0.8675214)(100%) = 86.752 %= **86.8%**
Mass percent of S = (0.1324786)(100%) = 13.248 %= **13%**

2.22 Plan: Dividing the mass of magnesium by the mass of the oxide gives the ratio. Multiply the mass of the second sample of magnesium oxide by this ratio to determine the mass of magnesium.

Solution:

a) If 1.25 g of MgO contains 0.754 g of Mg, then the mass ratio (or fraction) of magnesium in the oxide compound is $\frac{\text{mass Mg}}{\text{mass MgO}} = \frac{0.754 \text{ g Mg}}{1.25 \text{ g MgO}}$ = 0.6032 = **0.603**.

b) Mass (g) of magnesium = $\left(534 \text{ g MgO}\right)\left(\frac{0.6032 \text{ g Mg}}{1 \text{ g MgO}}\right)$ = 322.109 g = **322 g magnesium**

2.23 Plan: Dividing the mass of zinc by the mass of the sulfide gives the ratio. Multiply the mass of the second sample of zinc sulfide by this ratio to determine the mass of zinc.

Solution:

a) If 2.54 g of ZnS contains 1.70 g of Zn, then the mass ratio (or fraction) of zinc in the sulfide compound is

$$\frac{\text{mass Zn}}{\text{mass ZnS}} = \frac{1.70\text{ g Zn}}{2.54\text{ g ZnS}} = 0.66929 = \mathbf{0.669}.$$

b) Mass (g) of zinc = $\left(3.82\text{ kg ZnS}\right)\left(\frac{0.66929\text{ kg Zn}}{1\text{ kg ZnS}}\right)$ = 2.5567 kg= **2.56 kg zinc**

2.24 Plan: Since copper is a metal and sulfur is a nonmetal, the sample contains 88.39 g Cu and 44.61 g S. Calculate the mass fraction of each element in the sample by dividing the mass of element by the total mass of compound. Multiply the mass of the second sample of compound in grams by the mass fraction of each element to find the mass of each element in that sample.

Solution:

Mass (g) of compound = 88.39 g copper + 44.61 g sulfur = 133.00 g compound

Mass fraction of copper = $\left(\frac{88.39\text{ g copper}}{133.00\text{ g compound}}\right)$ = 0.664586

Mass (g) of copper = $\left(5264\text{ kg compound}\right)\left(\frac{10^3\text{ g compound}}{1\text{ kg compound}}\right)\left(\frac{0.664586\text{ g copper}}{1\text{ g compound}}\right)$

= 3.49838×10^6 g= **3.498×10^6 g copper**

Mass fraction of sulfur = $\left(\frac{44.61\text{ g sulfur}}{133.00\text{ g compound}}\right)$ = 0.335414

Mass (g) of sulfur = $\left(5264\text{ kg compound}\right)\left(\frac{10^3\text{ g compound}}{1\text{ kg compound}}\right)\left(\frac{0.335414\text{ g sulfur}}{1\text{ g compound}}\right)$

= 1.76562×10^6 g**= 1.766×10^6 g sulfur**

2.25 Plan: Since cesium is a metal and iodine is a nonmetal, the sample contains 63.94 g Cs and 61.06 g I. Calculate the mass fraction of each element in the sample by dividing the mass of element by the total mass of compound. Multiply the mass of the second sample of compound by the mass fraction of each element to find the mass of each element in that sample.

Solution:

Mass of compound = 63.94 g cesium + 61.06 g iodine = 125.00 g compound

Mass fraction of cesium = $\left(\frac{63.94\text{ g cesium}}{125.00\text{ g compound}}\right)$ = 0.51152

Mass (g) of cesium = $\left(38.77\text{ g compound}\right)\left(\frac{0.51152\text{ g cesium}}{1\text{ g compound}}\right)$ = 19.83163 g= **19.83 g cesium**

Mass fraction of iodine = $\left(\frac{61.06\text{ g iodine}}{125.00\text{ g compound}}\right)$ = 0.48848

Mass (g) of iodine = $\left(38.77\text{ g compound}\right)\left(\frac{0.48848\text{ g iodine}}{1\text{ g compound}}\right)$ = 18.9384 g= **18.94 g iodine**

2.26 Plan: The law of multiple proportions states that if two elements form two different compounds, the relative amounts of the elements in the two compounds form a whole-number ratio. To illustrate the law we must calculate the mass of one element to one gram of the other element for each compound and then compare this mass for the two compounds. The law states that the ratio of the two masses should be a small whole-number ratio such as 1:2, 3:2, 4:3, etc.

Solution:

Compound 1: $\frac{47.5 \text{ mass \% S}}{52.5 \text{ mass \% Cl}} = 0.90476 = 0.905$

Compound 2: $\frac{31.1 \text{ mass \% S}}{68.9 \text{ mass \% Cl}} = 0.451379 = 0.451$

Ratio: $\frac{0.905}{0.451} = 2.0067 = 2.00{:}1.00$

Thus, the ratio of the mass of sulfur per gram of chlorine in the two compounds is a small whole-number ratio of 2:1, which agrees with the law of multiple proportions.

2.27 Plan: The law of multiple proportions states that if two elements form two different compounds, the relative amounts of the elements in the two compounds form a whole-number ratio. To illustrate the law we must calculate the mass of one element to one gram of the other element for each compound and then compare this mass for the two compounds. The law states that the ratio of the two masses should be a small whole-number ratio such as 1:2, 3:2, 4:3, etc.

Solution:

Compound 1: $\frac{77.6 \text{ mass \% Xe}}{22.4 \text{ mass \% F}} = 3.4643 = 3.46$

Compound 2: $\frac{63.3 \text{ mass \% Xe}}{36.7 \text{ mass \% F}} = 1.7248 = 1.72$

Ratio: $\frac{3.46}{1.72} = 2.0116 = 2.01{:}1.00$

Thus, the ratio of the mass of xenon per gram of fluorine in the two compounds is a small whole-number ratio of 2:1, which agrees with the law of multiple proportions.

2.28 Plan: Calculate the mass percent of calcium in dolomite by dividing the mass of calcium by the mass of the sample and multiply by 100. Compare this mass percent to that in fluorite. The compound with the larger mass percent of calcium is the richer source of calcium.

Solution:

Mass percent calcium = $\frac{1.70 \text{ g calcium}}{7.81 \text{ g dolomite}} \times 100\% = 21.767\% =$ **21.8% Ca**

Fluorite (51.4%) is the richer source of calcium.

2.29 Plan: Determine the mass percent of sulfur in each sample by dividing the grams of sulfur in the sample by the total mass of the sample and multiplying by 100. The coal type with the smallest mass percent of sulfur has the smallest environmental impact.

Solution:

$$\text{Mass \% in Coal A} = \left(\frac{11.3 \text{ g sulfur}}{378 \text{ g sample}}\right)(100\%) = 2.9894\% = 2.99\% \text{ S (by mass)}$$

$$\text{Mass \% in Coal B} = \left(\frac{19.0 \text{ g sulfur}}{495 \text{ g sample}}\right)(100\%) = 3.8384\% = 3.84\% \text{ S (by mass)}$$

$$\text{Mass \% in Coal C} = \left(\frac{20.6 \text{ g sulfur}}{675 \text{ g sample}}\right)(100\%) = 3.0519\% = 3.05\% \text{ S (by mass)}$$

Coal A has the smallest environmental impact.

2.30 We now know that atoms of one element may change into atoms of another element. We also know that atoms of an element can have different masses (isotopes). Finally, we know that atoms are divisible into smaller particles. Based on the best available information in 1805, Dalton was correct. This model is still useful, since its essence (even if not its exact details) remains true today.

2.31 Plan: This question is based on the law of definite composition. If the compound contains the same types of atoms, they should combine in the same way to give the same mass percentages of each of the elements.
Solution:
Potassium nitrate is a compound composed of three elements — potassium, nitrogen, and oxygen — in a specific ratio. If the ratio of these elements changed, then the compound would be changed to a different compound, for example, to potassium nitrite, with different physical and chemical properties. Dalton postulated that atoms of an element are identical, regardless of whether that element is found in India or Italy. Dalton also postulated that compounds result from the chemical combination of specific ratios of different elements. Thus, Dalton's theory explains why potassium nitrate, a compound comprised of three different elements in a specific ratio, has the same chemical composition regardless of where it is mined or how it is synthesized.

2.32 Plan: Review the discussion of the experiments in this chapter.
Solution:
Millikan determined the minimum *charge* on an oil drop and that the minimum charge was equal to the charge on one electron. Using Thomson's value for the *mass/charge ratio* of the electron and the determined value for the charge on one electron, Millikan calculated the mass of an electron (charge/(charge/mass)) to be $9.109\text{x}10^{-28}$ g.

2.33 Plan: The charges on the oil droplets should be whole-number multiples of a minimum charge. Determine that minimum charge by dividing the charges by small integers to find the common factor.
Solution:

$$-3.204\text{x}10^{-19}\ \text{C}/2 = -1.602\text{x}10^{-19}\ \text{C}$$
$$-4.806\text{x}10^{-19}\ \text{C}/3 = -1.602\text{x}10^{-19}\ \text{C}$$
$$-8.010\text{x}10^{-19}\ \text{C}/5 = -1.602\text{x}10^{-19}\ \text{C}$$
$$-1.442\text{x}10^{-18}\ \text{C}/4 = -1.602\text{x}10^{-19}\ \text{C}$$

The value **$-1.602\text{x}10^{-19}$ C** is the common factor and is the charge for the electron.

2.34 Thomson's "plum pudding" model described the atom as a "blob" of positive charge with tiny electrons embedded in it. The electrons could be easily removed from the atoms when a current was applied and ejected as a stream of "cathode rays."

2.35 Rutherford and co-workers expected that the alpha particles would pass through the foil essentially unaffected, or perhaps slightly deflected or slowed down. The observed results (most passing through straight, a few deflected, a very few at large angles) were partially consistent with expectations, but the large-angle scattering could not be explained by Thomson's model. The change was that Rutherford envisioned a small (but massive) positively charged nucleus in the atom, capable of deflecting the alpha particles as observed.

2.36 Plan: Re-examine the definitions of atomic number and the mass number.
Solution:
The atomic number is the number of protons in the nucleus of an atom. When the atomic number changes, the identity of the element also changes. The mass number is the total number of protons and neutrons in the nucleus of an atom. Since the identity of an element is based on the number of protons and not the number of neutrons, the mass number can vary (by a change in number of neutrons) without changing the identity of the element.

2.37 Plan: Recall that the mass number is the sum of protons and neutrons while the atomic number is the number of protons.
Solution:
Mass number (protons plus neutrons) – atomic number (protons) = **number of neutrons (c)**.

2.38 The actual masses of the protons, neutrons, and electrons are not whole numbers so their sum is not a whole number.

2.39 Plan: The superscript is the mass number, the sum of the number of protons and neutrons. Consult the periodic table to get the atomic number (the number of protons). The mass number – the number of protons = the number of neutrons. For atoms, the number of protons and electrons are equal.

Solution:

Isotope	Mass Number	# of Protons	# of Neutrons	# of Electrons
^{36}Ar	36	18	18	18
^{38}Ar	38	18	20	18
^{40}Ar	40	18	22	18

2.40 Plan: The superscript is the mass number, the sum of the number of protons and neutrons. Consult the periodic table to get the atomic number (the number of protons). The mass number – the number of protons = the number of neutrons. For atoms, the number of protons and electrons are equal.

Solution:

Isotope	Mass Number	# of Protons	# of Neutrons	# of Electrons
^{35}Cl	35	17	18	17
^{37}Cl	37	17	20	17

2.41 Plan: The superscript is the mass number (A), the sum of the number of protons and neutrons; the subscript is the atomic number (Z, number of protons). The mass number – the number of protons = the number of neutrons. For atoms, the number of protons = the number of electrons.

Solution:

a) $^{16}_{8}O$ and $^{17}_{8}O$ have the **same number of protons and electrons** (8), but different numbers of neutrons.

$^{16}_{8}O$ and $^{17}_{8}O$ are isotopes of oxygen, and $^{16}_{8}O$ has 16 – 8 = 8 neutrons whereas $^{17}_{8}O$ has 17 – 8 = 9 neutrons.

Same *Z* value

b) $^{40}_{18}Ar$ and $^{41}_{19}K$ have the **same number of neutrons** (Ar: 40 – 18 = 22; K: 41 – 19 = 22) but different numbers of protons and electrons (Ar = 18 protons and 18 electrons; K = 19 protons and 19 electrons). **Same *N* value**

c) $^{60}_{27}Co$ and $^{60}_{28}Ni$ have different numbers of protons, neutrons, and electrons. Co: 27 protons, 27 electrons, and 60 – 27 = 33 neutrons; Ni: 28 protons, 28 electrons and 60 – 28 = 32 neutrons. However, both have a mass number of 60. **Same *A* value**

2.42 Plan: The superscript is the mass number (A), the sum of the number of protons and neutrons; the subscript is the atomic number (Z, number of protons). The mass number – the number of protons = the number of neutrons. For atoms, the number of protons = the number of electrons.

Solution:

a)) $^{3}_{1}H$ and $^{3}_{2}He$ have different numbers of protons, neutrons, and electrons. H: 1 proton, 1 electron, and 3 – 1 = 2 neutrons; He: 2 protons, 2 electrons, and 3 – 2 = 1 neutron. However, both have a mass number of 3.

Same *A* value

b) $^{14}_{6}C$ and $^{15}_{7}N$ have the **same number of neutrons** (C: 14 – 6 = 8; N: 15 – 7 = 8) but different numbers of protons and electrons (C = 6 protons and 6 electrons; N = 7 protons and 7 electrons). **Same *N* value**

c) $^{19}_{9}F$ and $^{18}_{9}F$ have the **same number of protons and electrons** (9), but different numbers of neutrons.

$^{19}_{9}F$ and $^{18}_{9}F$ are isotopes of oxygen, and $^{19}_{9}F$ has 19 – 9 = 10 neutrons whereas $^{18}_{9}F$ has 18 – 9 = 9 neutrons.

Same *Z* value

2.43 Plan: Combine the particles in the nucleus (protons + neutrons) to give the mass number (superscript, A). The number of protons gives the atomic number (subscript, Z) and identifies the element.

Solution:

a) A = 18 + 20 = 38; Z = 18; $\mathbf{^{38}_{18}Ar}$

b) A = 25 + 30 = 55; Z = 25; $\mathbf{^{55}_{25}Mn}$

c) A = 47 + 62 = 109; Z = 47; $\mathbf{^{109}_{47}Ag}$

2.44 Plan: Combine the particles in the nucleus (protons + neutrons) to give the mass number (superscript, *A*). The number of protons gives the atomic number (subscript, *Z*) and identifies the element.

Solution:

a) $A = 6 + 7 = 13$; $Z = 6$; $\mathbf{^{13}_{6}C}$

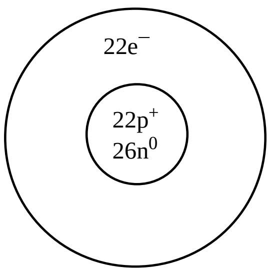
b) $A = 40 + 50 = 90$; $Z = 40$; $\mathbf{^{90}_{40}Zr}$

c) $A = 28 + 33 = 61$; $Z = 28$; $\mathbf{^{61}_{28}Ni}$

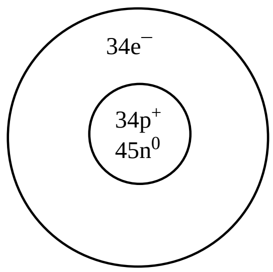
2.45 Plan: Determine the number of each type of particle. The superscript is the mass number (*A*) and the subscript is the atomic number (*Z*, number of protons). The mass number – the number of protons = the number of neutrons. For atoms, the number of protons = the number of electrons. The protons and neutrons are in the nucleus of the atom.

Solution:

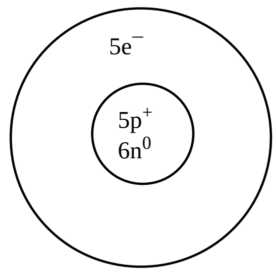
a) $^{48}_{22}Ti$
22 protons
22 electrons
48 – 22 = 26 neutrons



b) $^{79}_{34}Se$
34 protons
34 electrons
79 – 34 = 45 neutrons



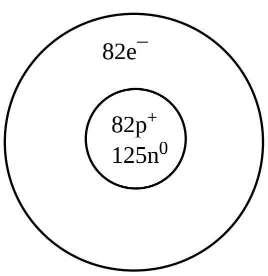
c) $^{11}_{5}B$
5 protons
5 electrons
11 – 5 = 6 neutrons



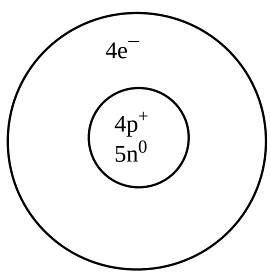
2.46 Plan: Determine the number of each type of particle. The superscript is the mass number (*A*) and the subscript is the atomic number (*Z*, number of protons). The mass number – the number of protons = the number of neutrons. For atoms, the number of protons = the number of electrons. The protons and neutrons are in the nucleus of the atom.

Solution:

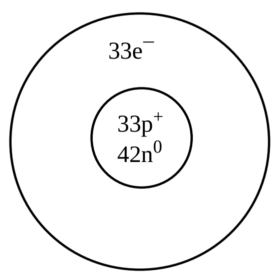
a) $^{207}_{82}Pb$
82 protons
82 electrons
207 – 82 = 125 neutrons



b) $^{9}_{4}Be$
4 protons
4 electrons
9 – 4 = 5 neutrons



c) $^{75}_{33}As$
33 protons
33 electrons
75 – 33 = 42 neutrons



2.47 Plan: To calculate the atomic mass of an element, take a weighted average based on the natural abundance of the isotopes: (isotopic mass of isotope 1 x fractional abundance) + (isotopic mass of isotope 2 x fractional abundance).

Solution:

$$\text{Atomic mass of gallium} = (68.9256\text{ u})\left(\frac{60.11\%}{100\%}\right) + (70.9247\text{ u})\left(\frac{39.89\%}{100\%}\right) = 69.7230\text{ u} = \mathbf{69.72\text{ u}}$$

2.48 Plan: To calculate the atomic mass of an element, take a weighted average based on the natural abundance of the isotopes: (isotopic mass of isotope 1 x fractional abundance) + (isotopic mass of isotope 2 x fractional abundance) + (isotopic mass of isotope 3 x fractional abundance).
Solution:

$$\text{Atomic mass of Mg} = (23.9850\ \text{u})\left(\frac{78.99\%}{100\%}\right)+(24.9858\ \text{u})\left(\frac{10.00\%}{100\%}\right)+(25.9826\ \text{u})\left(\frac{11.01\%}{100\%}\right)$$

$$= 24.3050\ \text{u} = \mathbf{24.31\ u}$$

2.49 Plan: To find the percent abundance of each Cl isotope, let x equal the fractional abundance of ^{35}Cl and (1 – x) equal the fractional abundance of ^{37}Cl since the sum of the fractional abundances must equal 1. Remember that atomic mass = (isotopic mass of ^{35}Cl x fractional abundance) + (isotopic mass of ^{37}Cl x fractional abundance).
Solution:
Atomic mass = (isotopic mass of ^{35}Cl x fractional abundance) + (isotopic mass of ^{37}Cl x fractional abundance)

35.4527 u = 34.9689 u(x) + 36.9659 u(1 – x)
35.4527 u = 34.9689 u(x) + 36.9659 u – 36.9659 u(x)
35.4527 u = 36.9659 u – 1.9970 u(x)
1.9970 u(x) = 1.5132 u
x = 0.75774 and 1 – x = 1 – 0.75774 = 0.24226
% abundance **^{35}Cl = 75.774%** % abundance **^{37}Cl = 24.226%**

2.50 Plan: To find the percent abundance of each Cu isotope, let x equal the fractional abundance of ^{63}Cu and (1 – x) equal the fractional abundance of ^{65}Cu since the sum of the fractional abundances must equal 1. Remember that atomic mass = (isotopic mass of ^{63}Cu x fractional abundance) + (isotopic mass of ^{65}Cu x fractional abundance).
Solution:
Atomic mass = (isotopic mass of ^{63}Cu x fractional abundance) + (isotopic mass of ^{65}Cu x fractional abundance)

63.546 u = 62.9396 u(x) + 64.9278 u(1 – x)
63.546 u = 62.9396 u(x) + 64.9278 u – 64.9278 u(x)
63.546 u = 64.9278 u – 1.9882 u(x)
1.9882 u(x) = 1.3818 u
x = 0.69500 and 1 – x = 1 – 0.69500 = 0.30500
% abundance **^{63}Cu = 69.50%** % abundance **^{65}Cu = 30.50%**

2.51 Iodine has more protons in its nucleus (higher *Z*), but iodine atoms must have, on average, fewer neutrons than Te atoms and thus a lower atomic mass.

2.52 Plan: Review the section in the chapter on the periodic table.
Solution:
a) In the modern periodic table, the elements are arranged in order of increasing atomic **number**.
b) Elements in a **column or group** (or family) have similar chemical properties, not those in the same period or row.
c) Elements can be classified as **metals**, metalloids, or nonmetals.

2.53 The **metalloids** lie along the "staircase" line, with properties intermediate between metals and nonmetals.

2.54 Plan: Review the section on the classification of elements as metals, nonmetals, or metalloids.
Solution:
To the left of the "staircase" are the metals, which are generally hard, shiny, malleable, ductile, good conductors of heat and electricity, and form positive ions by losing electrons. To the right of the "staircase" are the nonmetals, which are generally soft or gaseous, brittle, dull, poor conductors of heat and electricity, and form negative ions by gaining electrons.

2.55 Plan: Review the properties of these two columns in the periodic table.
Solution:
The alkali metals (Group 1) are metals and readily lose one electron to form cations whereas the halogens (Group 17) are nonmetals and readily gain one electron to form anions.

2.56 Plan: Locate each element on the periodic table. The Z value is the atomic number of the element. Metals are to the left of the "staircase," nonmetals are to the right of the "staircase," and the metalloids are the elements that lie along the "staircase" line.

Solution:

a) Germanium	Ge	14	metalloid
b) Phosphorus	P	15	nonmetal
c) Helium	He	18	nonmetal
d) Lithium	Li	1	metal
e) Molybdenum	Mo	6	metal

2.57 Plan: Locate each element on the periodic table. The Z value is the atomic number of the element. Metals are to the left of the "staircase," nonmetals are to the right of the "staircase," and the metalloids are the elements that lie along the "staircase" line.

Solution:

a) Arsenic	As	15	metalloid
b) Calcium	Ca	2	metal
c) Bromine	Br	17	nonmetal
d) Potassium	K	1	metal
e) Aluminum	Al	13	metal

2.58 Plan: Review the section in the chapter on the periodic table. Remember that alkaline earth metals are in Group 2, the halogens are in Group 17, and the metalloids are the elements that lie along the "staircase" line; periods are horizontal rows.

Solution:

a) The symbol and atomic number of the heaviest alkaline earth metal are **Ra** and **88**.

b) The symbol and atomic number of the lightest metalloid in Group 14 are **Si** and **14**.

c) The symbol and atomic mass of the coinage metal whose atoms have the fewest electrons are **Cu** and **63.55 u**.

d) The symbol and atomic mass of the halogen in Period 4 are **Br** and **79.90 u**.

2.59 Plan: Review the section in the chapter on the periodic table. Remember that the noble gases are in Group 18, the alkali metals are in Group 1, and the transition elements are the groups of elements located between Groups 2 and 13; periods are horizontal rows and metals are located to the left of the "staircase" line.

Solution:

a) The symbol and atomic number of the heaviest nonradioactive noble gas are **Xe** and **54**, respectively.

b) The symbol and group number of the Period 5 transition element whose atoms have the fewest protons are **Y** and **3**.

c) The symbol and atomic number of the only metallic chalcogen are **Po** and **84**.

d) The symbol and number of protons of the Period 4 alkali metal atom are **K** and **19**.

2.60 Plan: Review the section of the chapter on the formation of ionic compounds.

Solution:

Reactive metals and nometals will form **ionic** bonds, in which one or more electrons are transferred from the metal atom to the nonmetal atom to form a cation and an anion, respectively. The oppositely charged ions attract, forming the ionic bond.

2.61 Plan: Review the section of the chapter on the formation of covalent compounds.

Solution:

Two nonmetals will form **covalent** bonds, in which the atoms share two or more electrons.

2.62 The total positive charge of the cations is balanced by the total negative charge of the anions.

2.63 Plan: Assign charges to each of the ions. Since the sizes are similar, there are no differences due to the sizes.
Solution:
Coulomb's law states the energy of attraction in an ionic bond is directly proportional to the *product of charges* and inversely proportional to the *distance between charges*. The *product of charges* in MgO (+2 x –2 = –4) is greater than the *product of charges* in LiF (+1 x –1 = –1). Thus, **MgO** has stronger ionic bonding.

2.64 There are no molecules; BaF_2 is an ionic compound consisting of Ba^{2+} and F^- ions.

2.65 There are no ions present; P and O are both nonmetals, and they will bond covalently to form P_4O_6 molecules.

2.66 Plan: Locate these groups on the periodic table and assign charges to the ions that would form.
Solution:
The monatomic ions of Group 1 have a +1 charge (e.g., Li^+, Na^+, and K^+) whereas the monatomic ions of Group 17 have a –1 charge (e.g., F^-, Cl^-, and Br^-). Elements gain or lose electrons to form ions with the same number of electrons as the nearest noble gas. For example, Na loses one electron to form a cation with the same number of electrons as Ne. The halogen F gains one electron to form an anion with the same number of electrons as Ne.

2.67 Plan: A metal and a nonmetal will form an ionic compound. Locate these elements on the periodic table and predict their charges.
Solution:
Magnesium chloride ($MgCl_2$) is an ionic compound formed from a metal (magnesium) and a nonmetal (chlorine). Magnesium atoms transfer electrons to chlorine atoms. Each magnesium atom loses two electrons to form a Mg^{2+} ion and the same number of electrons (10) as the noble gas neon. Each chlorine atom gains one electron to form a Cl^- ion and the same number of electrons (18) as the noble gas argon. The Mg^{2+} and Cl^- ions attract each other to form an ionic compound with the ratio of one Mg^{2+} ion to two Cl^- ions. The total number of electrons lost by the magnesium atoms equals the total number of electrons gained by the chlorine atoms.

2.68 Plan: A metal and a nonmetal will form an ionic compound. Locate these elements on the periodic table and predict their charges.
Solution:
Potassium sulfide (K_2S) is an ionic compound formed from a metal (potassium) and a nonmetal (sulfur). Potassium atoms transfer electrons to sulfur atoms. Each potassium atom loses one electron to form an ion with +1 charge and the same number of electrons (18) as the noble gas argon. Each sulfur atom gains two electrons to form an ion with a –2 charge and the same number of electrons (18) as the noble gas argon. The oppositely charged ions, K^+ and S^{2-}, attract each other to form an ionic compound with the ratio of two K^+ ions to one S^{2-} ion. The total number of electrons lost by the potassium atoms equals the total number of electrons gained by the sulfur atoms.

2.69 Plan: Recall that ionic bonds occur between metals and nonmetals, whereas covalent bonds occur between nonmetals.
Solution:
KNO_3 shows **both** ionic and covalent bonding, covalent bonding between the N and O in NO_3^- and ionic bonding between the NO_3^- and the K^+.

2.70 Plan: Locate these elements on the periodic table and predict what ions they will form. For cations (metals), ion charge = group number; for anions (nonmetals), ion charge = group number minus 18 .
Solution:
Potassium (K) is in Group **1** and forms the $\mathbf{K^+}$ ion. Iodine (I) is in Group**17** and forms the $\mathbf{I^-}$ ion (17 –18 = –1).

2.71 Plan: Locate these elements on the periodic table and predict what ions they will form. For cations (metals), ion charge = group number; for anions (nonmetals), ion charge = group number minus 18 .
Solution:
Barium in Group **2** forms a +2 ion: $\mathbf{Ba^{2+}}$. Selenium in Group **16** forms a –2 ion: $\mathbf{Se^{2-}}$ (16 –18 = –2).

2.72 Plan: Use the number of protons (atomic number) to identify the element. Add the number of protons and neutrons together to get the mass number. Locate the element on the periodic table and assign its group and period number.
Solution:

a) Oxygen (atomic number = 8)	mass number = 8p + 9n = 17	Group 16	Period 2
b) Fluorine (atomic number = 9)	mass number = 9p + 10n = 19	Group 17	Period 2
c) Calcium (atomic number = 20)	mass number = 20p + 20n = 40	Group 2	Period 4

2.73 Plan: Use the number of protons (atomic number) to identify the element. Add the number of protons and neutrons together to get the mass number. Locate the element on the periodic table and assign its group and period number.
Solution:

a) Bromine (atomic number = 35)	mass number = 35p + 44n = 79	Group 17	Period 4
b) Nitrogen (atomic number = 7)	mass number = 7p + 8n = 15	Group 15	Period 2
c) Rubidium (atomic number = 37)	mass number = 37p + 48n = 85	Group 1	Period 5

2.74 Plan: Determine the charges of the ions based on their position on the periodic table. For cations (metals), ion charge = group number; for anions (nonmetals), ion charge = group number minus 18. Next, determine the ratio of the charges to get the ratio of the ions.
Solution:
Lithium [Group 1] forms the Li^+ ion; oxygen [Group 16] forms the O^{2-} ion (16 -18 = –2). The ionic compound that forms from the combination of these two ions must be electrically neutral, so two Li^+ ions combine with one O^{2-} ion to form the compound Li_2O. There are twice as many Li^+ ions as O^{2-} ions in a sample of Li_2O.

$$\text{Number of } O^{2-} \text{ ions} = (8.4\text{x}10^{21}\ Li^+ \text{ ions})\left(\frac{1\ O^{2-}\text{ ion}}{2\ Li^+\text{ ions}}\right) = \mathbf{4.2x10^{21}\ O^{2-}\ ions}$$

2.75 Plan: Determine the charges of the ions based on their position on the periodic table. For cations (metals), ion charge = group number; for anions (nonmetals), ion charge = group number minus 18. Next, determine the ratio of the charges to get the ratio of the ions.
Solution:
Ca [Group 2] forms Ca^{2+} and I [Group 17] forms I^- ions (17 -18 = –1). The ionic compound that forms from the combination of these two ions must be electrically neutral, so one Ca^{2+} ion combines with two I^- ions to form the compound CaI_2. There are twice as many I^- ions as Ca^{2+} ions in a sample of CaI_2.

$$\text{Number of } I^- \text{ ions} = (7.4\text{x}10^{21}\ Ca^{2+} \text{ ions})\left(\frac{2\ I^-\text{ ions}}{1\ Ca^{2+}\text{ ion}}\right) = 1.48\text{x}10^{22}\text{ ions} = \mathbf{1.5x10^{22}\ I^-\ ions}$$

2.76 Plan: The key is the size of the two alkali metal ions. The charges on the sodium and potassium ions are the same as both are in Group 1, so there will be no difference due to the charge. The chloride ions are the same in size and charge, so there will be no difference due to the chloride ion.
Solution:
Coulomb's law states that the energy of attraction in an ionic bond is directly proportional to the *product of charges* and inversely proportional to the *distance between charges*. The *product of the charges* is the same in both compounds because both sodium and potassium ions have a +1 charge. Attraction increases as distance decreases, so the ion with the smaller radius, Na^+, will form a stronger ionic interaction (**NaCl**).

2.77 Plan: The key is the charge of the two metal ions. The sizes of the lithium and magnesium ions are about the same (magnesium is slightly smaller), so there will be little difference due to ion size. The oxide ions are the same in size and charge, so there will be no difference due to the oxide ion.
Solution:
Coulomb's law states the energy of attraction in an ionic bond is directly proportional to the *product of charges* and inversely proportional to the *distance between charges*. The *product of charges* in MgO (+2 x –2 = –4) is greater than the *product of charges* in Li_2O (+1 x –2 = –2). Thus, **MgO** has stronger ionic bonding.

2.78 Plan: Review the definition of molecular formula.
Solution:
The subscripts in the formula, MgF_2, give the number of ions in a formula unit of the ionic compound. The subscripts indicate that there are two F^- ions for every one Mg^{2+} ion. Using this information and the mass of each element, we could calculate the percent mass of each element.

2.79 Plan: Review the definitions of molecular and structural formulas.
Solution:
Both the structural and molecular formulas show the actual numbers of the atoms of the molecule; in addition, the structural formula shows the arrangement of the atoms (i.e., how the atoms are connected to each other).

2.80 Plan: Review the concepts of atoms and molecules.
Solution:
The mixture is similar to the sample of hydrogen peroxide in that both contain 20 billion oxygen atoms and 20 billion hydrogen atoms since both O_2 and H_2O_2 contain 2 oxygen atoms per molecule and both H_2 and H_2O_2 contain 2 hydrogen atoms per molecule. They differ in that they contain different types of molecules: H_2O_2 molecules in the hydrogen peroxide sample and H_2 and O_2 molecules in the mixture. In addition, the mixture contains 20 billion molecules (10 billion H_2 molecules + 10 billion O_2 molecules) while the hydrogen peroxide sample contains 10 billion molecules.

2.81 Plan: Review the rules for naming compounds.
Solution:
Roman numerals are used when naming ionic compounds that contain a metal that can form more than one ion. This is generally true for the transition metals, but it can be true for some non-transition metals as well (e.g., Sn).

2.82 Plan: Review the rules for naming compounds.
Solution:
Greek prefixes are used only in naming covalent compounds.

2.83 Molecular formulas cannot be written for ionic compounds since they only have ions and there are no molecules.

2.84 Plan: Locate each of the individual elements on the periodic table, and assign charges to each of the ions. For cations (metals), ion charge = group number; for anions (nonmetals), ion charge = group number minus 18. Find the smallest number of each ion that gives a neutral compound. To name ionic compounds with metals that form only one ion, name the metal, followed by the nonmetal name with an -ide suffix.
Solution:
a) Sodium is a metal that forms a +1 (Group **1**) ion and nitrogen is a nonmetal that forms a –3 ion (Group 1**5**, 15-18 = –3).

$$\overset{+1\ \ -3}{\mathrm{Na\ \ N}} \qquad \overset{+3\ \ -3}{\overset{+1}{\mathrm{Na_3N}}}$$

The compound is **Na_3N, sodium nitride**.

b) Oxygen is a nonmetal that forms a –2 ion (Group 1**6**, 16 -18 = –2) and strontium is a metal that forms a +2 ion (Group **2**).

$$\overset{+2\ -2}{\mathrm{Sr\ O}}$$

The compound is **SrO, strontium oxide**.

c) Aluminum is a metal that forms a +3 ion (Group **3**) and chlorine is a nonmetal that forms a –1 ion (Group 1**7**, 17-18 = –1).

$$\overset{+3\ \ -1}{\mathrm{Al\ Cl}} \qquad \overset{+3\ -3}{\overset{+3\ \ -1}{\mathrm{AlCl_3}}}$$

The compound is **$AlCl_3$, aluminum chloride**.

2.85 Plan: Locate each of the individual elements on the periodic table, and assign charges to each of the ions. For cations (metals), ion charge = group number; for anions (nonmetals), ion charge = group number minus 18. Find the smallest number of each ion that gives a neutral compound. To name ionic compounds with metals that form only one ion, name the metal, followed by the nonmetal name with an -ide suffix.

Solution:
a) Cesium is a metal that forms a +1 (Group **1**) ion and bromine is a nonmetal that forms a –1 ion (Group 1**7**, 17 -18 = –1).

+1 –1
Cs Br The compound is **CsBr, cesium bromide**.

b) Sulfur is a nonmetal that forms a –2 ion (Group 1**6**, 16-18 = –2) and barium is a metal that forms a +2 ion (Group **2**).

+2 –2
Ba S The compound is **BaS, barium sulfide**.

c) Fluorine is a nonmetal that forms a –1 ion (Group 1**7**, 17-18 = –1) and calcium is a metal that forms a +2 ion (Group **2**).

+2 –1 → +2 –1 (–2)
Ca F → CaF_2 The compound is **CaF_2, calcium fluoride**.

2.86 Plan: Based on the atomic numbers (the subscripts) locate the elements on the periodic table. Once the atomic numbers are located, identify the element and based on its position, assign a charge. For cations (metals), ion charge = group number; for anions (nonmetals), ion charge = group number minus 18. Find the smallest number of each ion that gives a neutral compound. To name ionic compounds with metals that form only one ion, name the metal, followed by the nonmetal name with an -ide suffix.
Solution:
a) ${}_{12}L$ is the element Mg (Z = 12). Magnesium [Group **2**] forms the Mg^{2+} ion. ${}_{9}M$ is the element F (Z = 9). Fluorine [Group 17] forms the F^{-} ion (17-18 = –1). The compound formed by the combination of these two elements is **MgF_2, magnesium fluoride**.
b) ${}_{30}L$ is the element Zn (Z = 30). Zinc forms the Zn^{2+} ion (see Table 2.3). ${}_{16}M$ is the element S (Z = 16). Sulfur [Group 16] will form the S^{2-} ion (16-18 = –2). The compound formed by the combination of these two elements is **ZnS, zinc sulfide**.
c) ${}_{17}L$ is the element Cl (Z = 17). Chlorine [Group 17] forms the Cl^{-} ion (17-18 = –1). ${}_{38}M$ is the element Sr (Z = 38). Strontium [Group 2] forms the Sr^{2+} ion. The compound formed by the combination of these two elements is **$SrCl_2$, strontium chloride**.

2.87 Plan: Based on the atomic numbers (the subscripts) locate the elements on the periodic table. Once the atomic numbers are located, identify the element and based on its position, assign a charge. For (metals), ion charge = group number; for anions (nonmetals), ion charge = group number minus 18. Find the smallest number of each ion that gives a neutral compound. To name ionic compounds with metals that form only one ion, name the metal, followed by the nonmetal name with an -ide suffix.
Solution:
a) ${}_{37}Q$ is the element Rb (Z = 37). Rubidium [Group 1] forms the Rb^{+} ion. ${}_{35}R$ is the element Br (Z = 35). Bromine [Group 17] forms the Br^{-} ion (17-18 = –1). The compound formed by the combination of these two elements is **RbBr, rubidium bromide**.
b) ${}_{8}Q$ is the O (Z = 8). Oxygen [Group 16] will form the O^{2-} ion (16 –-18 = –2). ${}_{13}R$ is the element Al (Z = 13). Aluminum [Group13) forms the Al^{3+} ion. The compound formed by the combination of these two elements is **Al_2O_3, aluminum oxide**.
c) ${}_{20}Q$ is the element Ca (Z = 20). Calcium [Group 2] forms the Ca^{2+} ion. ${}_{53}R$ is the element I (Z = 53). Iodine [Group 17] forms the I^{-} ion (17 –18 = –1). The compound formed by the combination of these two elements is **CaI_2, calcium iodide**.

2.88 Plan: Review the rules for nomenclature covered in the chapter. For ionic compounds, name the metal, followed by the nonmetal name with an -ide suffix. For metals, like many transition metals, that can form more than one ion each with a different charge, the ionic charge of the metal ion is indicated by a Roman numeral within parentheses immediately following the metal's name.
Solution:
a) tin(IV) chloride = **$SnCl_4$** The (IV) indicates that the metal ion is Sn^{4+} which requires 4 Cl^{-} ions for a neutral compound.
b) $FeBr_3$ = **iron(III) bromide** (common name is ferric bromide); the charge on the iron ion is +3 to match the –3 charge of 3 Br^{-} ions. The +3 charge of the Fe is indicated by (III).
c) cuprous bromide = **CuBr** (cuprous is +1 copper ion, cupric is +2 copper ion).
d) Mn_2O_3 = **manganese(III) oxide** Use (III) to indicate the +3 ionic charge of Mn:

+6 –6
+3 –2
Mn_2O_3

2.89 Plan: Review the rules for nomenclature covered in the chapter. For ionic compounds containing polyatomic ions, name the metal, followed by the name of the polyatomic ion. Hydrates, compounds with a specific number of water molecules associated with them, are named with a prefix before the word hydrate to indicate the number of water molecules.
Solution:
a) Na_2HPO_4 **= sodium hydrogen phosphate** Sodium [Group 1] forms the Na^+ ion; HPO_4^{2-} is the hydrogen phosphate ion.
b) potassium carbonate dihydrate = $\mathbf{K_2CO_3{\bullet}2H_2O}$ Potassium [Group 1] forms the K^+ ion; carbonate is the CO_3^{2-} ion. Two K^+ ions are required to match the –2 charge of the carbonate ion. Dihydrate indicates two water molecules (“waters of hydration”) that are written after a centered dot.
c) $NaNO_2$ = **sodium nitrite** NO_2^- is the nitrite polyatomic ion.
d) ammonium perchlorate = $\mathbf{NH_4ClO_4}$ Ammonium is the polyatomic ion NH_4^+ and perchlorate is the polyatomic ion ClO_4^-. One NH_4^+ is required for every one ClO_4^- ion.

2.90 Plan: Review the rules for nomenclature covered in the chapter. For ionic compounds, name the metal, followed by the nonmetal name with an -ide suffix. For metals, like many transition metals, that can form more than one ion each with a different charge, the ionic charge of the metal ion is indicated by a Roman numeral within parentheses immediately following the metal’s name. Hydrates, compounds with a specific number of water molecules associated with them, are named with a prefix before the word hydrate to indicate the number of water molecules.
Solution:
a) **cobalt(II) oxide** Cobalt forms more than one monatomic ion so the ionic charge must be indicated with a Roman numeral. Since the Co is paired with one O^{2-} ion, the charge of Co is +2.
b) $\mathbf{Hg_2Cl_2}$ The Roman numeral I indicates that mercury has an ionic charge of +1; mercury is an unusual case in which the +1 ion formed is Hg_2^{2+}, not Hg^+.
c) **lead(II) acetate trihydrate** The $C_2H_3O_2^-$ ion has a –1 charge (see Table 2.5); since there are two of these ions, the lead ion has a +2 charge which must be indicated with the Roman numeral II. The $\bullet 3H_2O$ indicates a hydrate in which the number of H_2O molecules is indicated by the prefix tri-.
d) $\mathbf{Cr_2O_3}$ “chromic” denotes a +3 charge (see Table 2.4), oxygen has a –2 charge: $\overset{+3\ -2}{CrO} \rightarrow \overset{+6\ -6}{Cr_2O_3}$

2.91 Plan: Review the rules for nomenclature covered in the chapter. For ionic compounds containing polyatomic ions, name the metal, followed by the name of the polyatomic ion. For metals, like many transition metals, that can form more than one ion each with a different charge, the ionic charge of the metal ion is indicated by a Roman numeral within parentheses immediately following the metal’s name.
Solution:
a) **tin(IV) sulfite** Tin forms more than one monatomic ion so the ionic charge must be indicated with a Roman numeral. Each SO_3^{2-} polyatomic ion has a charge of –2, so the ionic charge of tin is +4.
b) $\mathbf{K_2Cr_2O_7}$ Dichromate is the polyatomic ion $Cr_2O_7^{2-}$; two K^+ ions are required for a neutral compound.
c) **iron(II) carbonate** Iron forms more than one monatomic ion so the ionic charge must be indicated with a Roman numeral. The CO_3^{2-} polyatomic ion has a charge of –2, so the ionic charge of iron is +2.
d) $\mathbf{Cu(NO_3)_2}$ The Roman numeral II indicates that copper has an ionic charge of +2; two NO_3^- polyatomic ions are required for a neutral compound.

2.92 Plan: Review the rules for nomenclature covered in the chapter. For metals, like many transition metals, that can form more than one ion each with a different charge, the ionic charge of the metal ion is indicated by a Roman numeral within parentheses immediately following the metal’s name. Compounds must be neutral.
Solution:
a) Barium [Group 2] forms Ba^{2+} and oxygen [Group 16] forms O^{2-} (16 -18 = –2) so the neutral compound forms from one Ba^{2+} ion and one O^{2-} ion. Correct formula is **BaO**.
b) Iron(II) indicates Fe^{2+} and nitrate is NO_3^- so the neutral compound forms from one iron(II) ion and two nitrate ions. Correct formula is $\mathbf{Fe(NO_3)_2}$.
c) Mn is the symbol for manganese. Mg is the correct symbol for magnesium. Correct formula is **MgS**. Sulfide is the S^{2-} ion and sulfite is the SO_3^{2-} ion.

2.93 Plan: Review the rules for nomenclature covered in the chapter. For metals, like many transition metals, that can form more than one ion each with a different charge, the ionic charge of the metal ion is indicated by a Roman numeral within parentheses immediately following the metal's name. Compounds must be neutral.
Solution:
a) **copper(I) iodide** Cu is copper, not cobalt; since iodide is I^-, this must be copper(I).
b) **iron(III) hydrogen sulfate** HSO_4^- is hydrogen sulfate, and this must be iron(III) to be neutral.
c) **magnesium dichromate** Mg forms Mg^{2+} and $Cr_2O_7^{2-}$ is named dichromate ion.

2.94 Plan: Acids donate H^+ ion to the solution, so the acid is a combination of H^+ and a negatively charged ion. Binary acids (H plus one other nonmetal) are named hydro- + nonmetal root + -ic acid. Oxoacids (H + an oxoanion) are named by changing the suffix of the oxoanion: -ate becomes -ic acid and -ite becomes -ous acid.
Solution:
a) Hydrogen sulfate is HSO_4^-, so its source acid is **H_2SO_4**. Name of acid is **sulfuric acid** (-ate becomes -ic acid).
b) **HIO_3, iodic acid** IO_3^- is the iodate ion: -ate becomes -ic acid.
c) Cyanide is CN^-; its source acid is **HCN hydrocyanic acid** (binary acid).
d) **H_2S, hydrosulfuric acid** (binary acid).

2.95 Plan: Acids donate H^+ ion to the solution, so the acid is a combination of H^+ and a negatively charged ion. Binary acids (H plus one other nonmetal) are named hydro- + nonmetal root + -ic acid. Oxoacids (H + an oxoanion) are named by changing the suffix of the oxoanion: -ate becomes -ic acid and -ite becomes -ous acid.
Solution:
a) Perchlorate is ClO_4^-, so the source acid is **$HClO_4$**. Name of acid is **perchloric acid** (-ate becomes -ic acid).
b) **nitric acid, HNO_3** NO_3^- is the nitrate ion: -ate becomes -ic acid.
c) Bromite is BrO_2^-, so the source acid is **$HBrO_2$**. Name of acid is **bromous acid** (-ite becomes -ous acid).
d) **hydrofluoric acid, HF** (binary acid)

2.96 Plan: Use the formulas of the polyatomic ions. Recall that oxoacids are named by changing the suffix of the oxoanion: -ate becomes -ic acid and -ite becomes -ous acid. Compounds must be neutral.
Solution:
a) ammonium ion = **NH_4^+** ammonia = **NH_3**
b) magnesium sulfide = **MgS** magnesium sulfite = **$MgSO_3$** magnesium sulfate = **$MgSO_4$**
Sulfide = S^{2-}; sulfite = SO_3^{2-}; sulfate = SO_4^{2-}.
c) hydrochloric acid = **HCl** chloric acid = **$HClO_3$** chlorous acid = **$HClO_2$**
Binary acids (H plus one other nonmetal) are named hydro- + nonmetal root + -ic acid. Chloric indicates the polyatomic ion ClO_3^- while chlorous indicates the polyatomic ion ClO_2^-.
d) cuprous bromide = **CuBr** cupric bromide = **$CuBr_2$**
The suffix -ous indicates the lower charge, +1, while the suffix -ic indicates the higher charge, +2.

2.97 Plan: Use the formulas of the polyatomic ions. For metals, like many transition metals, that can form more than one ion each with a different charge, the ionic charge of the metal ion is indicated by a Roman numeral within parentheses immediately following the metal's name. Compounds must be neutral.
Solution:
a) lead(II) oxide = **PbO** lead(IV) oxide = **PbO_2**
Lead(II) indicates Pb^{2+} while lead(IV) indicates Pb^{4+}.
b) lithium nitride = **Li_3N** lithium nitrite = **$LiNO_2$** lithium nitrate = **$LiNO_3$**
Nitride = N^{3-}; nitrite = NO_2^-; nitrate = NO_3^-.
c) strontium hydride = **SrH_2** strontium hydroxide = **$Sr(OH)_2$**
Hydride = H^-; hydroxide = OH^-.
d) magnesium oxide = **MgO** manganese(II) oxide = **MnO**

2.98 Plan: This compound is composed of two nonmetals. The element with the lower group number is named first. Greek numerical prefixes are used to indicate the number of atoms of each element in the compound.
Solution:
disulfur tetrafluoride **S_2F_4** Di- indicates two S atoms and tetra- indicates four F atoms.

2.99 Plan: This compound is composed of two nonmetals. When a compound contains oxygen and a halogen, the halogen is named first. Greek numerical prefixes are used to indicate the number of atoms of each element in the compound.

Solution:

dichlorine monoxide **Cl_2O** Di- indicates two Cl atoms and mono- indicates one O atom.

2.100 Plan: Review the nomenclature rules in the chapter. For ionic compounds, name the metal, followed by the nonmetal name with an -ide suffix. For metals, like many transition metals, that can form more than one ion each with a different charge, the ionic charge of the metal ion is indicated by a Roman numeral within parentheses immediately following the metal's name. Binary acids (H plus one other nonmetal) are named hydro- + nonmetal root + -ic acid.

Solution:

a) Calcium(II) dichloride, $CaCl_2$: The name becomes **calcium chloride** because calcium does not require "(II)" since it only forms +2 ions. Prefixes like di- are only used in naming covalent compounds between nonmetal elements.

b) Copper(II) oxide, Cu_2O: The charge on the oxide ion is O^{2-}, which makes each copper a Cu^+. The name becomes **copper(I) oxide** to match the charge on the copper.

c) Stannous fluoride, SnF_4: Stannous refers to Sn^{2+}, but the tin in this compound is Sn^{4+} due to the charge on the fluoride ion. The tin(IV) ion is the stannic ion; this gives the name **stannic fluoride or tin(IV) fluoride**.

d) Hydrogen chloride acid, HCl: Binary acids consist of the root name of the nonmetal (chlor in this case) with a hydro- prefix and an -ic suffix. The word acid is also needed. This gives the name **hydrochloric acid**.

2.101 Plan: Review the nomenclature rules in the chapter. For ionic compounds, name the metal, followed by the nonmetal name with an -ide suffix. For metals, like many transition metals, that can form more than one ion each with a different charge, the ionic charge of the metal ion is indicated by a Roman numeral within parentheses immediately following the metal's name. Oxoacids (H + an oxoanion) are named by changing the suffix of the oxoanion: -ate becomes -ic acid and -ite becomes -ous acid. Greek numerical prefixes are used to indicate the number of atoms of each element in a compound composed of two nonmetals.

Solution:

a) Iron(III) oxide, Fe_3O_4: Iron(III) is Fe^{3+}, which combines with O^{2-} to give **Fe_2O_3**.

b) Chloric acid, HCl: HCl is hydrochloric acid. Chloric acid includes oxygen, and has the formula **$HClO_3$**.

c) Mercuric oxide, Hg_2O: The compound shown is mercurous oxide. Mercuric oxide contains Hg^{2+}, which combines with O^{2-} to give **HgO**.

d) Dichlorine heptoxide, Cl_2O_6: Heptoxide refers to seven, not six, oxygen atoms. The formula should be **Cl_2O_7**.

2.102 Plan: Break down each formula to the individual elements and count the number of atoms of each element by observing the subscripts. The molecular (formula) mass is the sum of the atomic masses of all of the atoms.

Solution:

a) There are **12 atoms of oxygen** in $Al_2(SO_4)_3$. The molecular mass is:

Al	=	2(26.98 u)	=	53.96 u
S	=	3(32.07 u)	=	96.21 u
O	=	12(16.00 u)	=	192.0 u
				342.2 u

b) There are **9 atoms of hydrogen** in $(NH_4)_2HPO_4$. The molecular mass is:

N	=	2(14.01 u)	=	28.02 u
H	=	9(1.008 u)	=	9.072 u
P	=	1(30.97u)	=	30.97 u
O	=	4(16.00 u)	=	64.00 u
				132.06 u

c) There are **8 atoms of oxygen** in $Cu_3(OH)_2(CO_3)_2$. The molecular mass is:

Cu	=	3(63.55 u)	=	190.6 u
O	=	8(16.00 u)	=	128.0 u
H	=	2(1.008 u)	=	2.016 u
C	=	2(12.01 u)	=	24.02 u
				344.6 u

2.103 Plan: Break down each formula to the individual elements and count the number of atoms of each element by observing the subscripts. The molecular (formula) mass is the sum of the atomic masses of all of the atoms.
Solution:
a) There are **9 atoms of hydrogen** in $C_6H_5COONH_4$. The molecular mass is:

C	=	7(12.01 u)	=	84.07 u
H	=	9(1.008 u)	=	9.072 u
O	=	2(16.00 u)	=	32.00 u
N	=	1(14.01 u)	=	14.01 u
				139.15 u

b) There are **2 atoms of nitrogen** in $N_2H_6SO_4$. The molecular mass is:

N	=	2(14.01 u)	=	28.02 u
H	=	6(1.008 u)	=	6.048 u
S	=	1(32.07 u)	=	32.07 u
O	=	4(16.00 u)	=	64.00 u
				130.14 u

c) There are 1**2 atoms of oxygen** in $Pb_4SO_4(CO_3)_2(OH)_2$. The molecular mass is:

Pb	=	4(207.2 u)	=	828.8 u
S	=	1(32.07 u)	=	32.07 u
O	=	12(16.00 u)	=	192.00 u
C	=	2(12.01 u)	=	24.02 u
H	=	2(1.008 u)	=	2.016 u
				1078.9 u

2.104 Plan: Review the rules for nomenclature covered in the chapter. For ionic compounds containing polyatomic ions, name the metal, followed by the name of the polyatomic ion. The molecular (formula) mass is the sum of the atomic masses of all of the atoms.
Solution:
a) **$(NH_4)_2SO_4$** ammonium is NH_4^+ and sulfate is SO_4^{2-}

N	=	2(14.01 u)	=	28.02 u
H	=	8(1.008 u)	=	8.064 u
S	=	1(32.07 u)	=	32.07 u
O	=	4(16.00 u)	=	64.00 u
				132.15 u

b) **NaH_2PO_4** sodium is Na^+ and dihydrogen phosphate is $H_2PO_4^-$

Na	=	1(22.99 u)	=	22.99 u
H	=	2(1.008 u)	=	2.016 u
P	=	1(30.97 u)	=	30.97 u
O	=	4(16.00 u)	=	64.00 u
				119.98 u

c) **$KHCO_3$** potassium is K^+ and bicarbonate is HCO_3^-

K	=	1(39.10 u)	=	39.10 u
H	=	1(1.008 u)	=	1.008 u
C	=	1(12.01 u)	=	12.01 u
O	=	3(16.00 u)	=	48.00 u
				100.12 u

2.105 Plan: Review the rules for nomenclature covered in the chapter. For ionic compounds containing polyatomic ions, name the metal, followed by the name of the polyatomic ion. The molecular (formula) mass is the sum of the atomic masses of all of the atoms.
Solution:
a) **$Na_2Cr_2O_7$** sodium is Na^+ and dichromate is $Cr_2O_7^{2-}$

Na	=	2(22.99 u)	=	45.98 u
Cr	=	2(52.00 u)	=	104.00 u
O	=	7(16.00 u)	=	112.00 u
				261.98 u

b) $\mathbf{NH_4ClO_4}$ ammonium is NH_4^+ and perchlorate is ClO_4^-

N	=	1(14.01 u)	=	14.01 u
H	=	4(1.008 u)	=	4.032 u
Cl	=	1(35.45 u)	=	35.45 u
O	=	4(16.00 u)	=	64.00 u
				117.49 u

c) $\mathbf{Mg(NO_2)_2 \bullet 3H_2O}$ magnesium is Mg^{2+}, nitrite is NO_2^-, and trihydrate is $3H_2O$

Mg	=	1(24.31 u)	=	24.31 u
N	=	2(14.01 u)	=	28.02 u
H	=	6(1.008 u)	=	6.048 u
O	=	7(16.00 u)	=	112.00 u
				170.38 u

2.106 Plan: Convert the names to the appropriate chemical formulas. The molecular (formula) mass is the sum of the masses of each atom times its atomic mass.

Solution:

a) dinitrogen pentoxide N_2O_5 (di- = 2 and penta- = 5)

N	=	2(14.01 u)	=	28.02 u
O	=	5(16.00 u)	=	80.00 u
				108.02 u

b) lead(II) nitrate $Pb(NO_3)_2$ (lead(II) is Pb^{2+} and nitrate is NO_3^-)

Pb	=	1(207.2 u)	=	207.2 u
N	=	2(14.01 u)	=	28.02 u
O	=	6(16.00 u)	=	96.00 u
				331.2 u

c) calcium peroxide CaO_2 (calcium is Ca^{2+} and peroxide is O_2^{2-})

Ca	=	1(40.08 u)	=	40.08 u
O	=	2(16.00 u)	=	32.00 u
				72.08 u

2.107 Plan: Convert the names to the appropriate chemical formulas. The molecular (formula) mass is the sum of the masses of each atom times its atomic mass.

Solution:

a) iron(II) acetate tetrahydrate $Fe(C_2H_3O_2)_2 \bullet 4H_2O$ (iron(II) is Fe^{2+}, acetate is $C_2H_3O_2^-$, and tetrahydrate is $4H_2O$)

Fe	=	1(55.85 u)	=	55.85 u
C	=	4(12.01 u)	=	48.04 u
H	=	14(1.008 u)	=	14.112 u
O	=	8(16.00 u)	=	128.00 u
				246.00 u

b) sulfur tetrachloride SCl_4 (tetra- = 4)

S	=	1 (32.07 u)	=	32.07 u
Cl	=	4(35.45 u)	=	141.80 u
				173.87 u

c) potassium permanganate $KMnO_4$ (potassium is K^+ and permanganate is MnO_4^-)

K	=	1(39.10 u)	=	39.10 u
Mn	=	1(54.94 u)	=	54.94 u
O	=	4(16.00 u)	=	64.00 u
				158.04 u

2.108 Plan: Use the chemical symbols and count the atoms of each type to give a molecular formula. Use the nomenclature rules in the chapter to derive the name. The molecular (formula) mass is the sum of the masses of each atom times its atomic mass.

Solution:
a) Formula is $\mathbf{SO_3}$. Name is **sulfur trioxide** (the prefix tri- indicates 3 oxygen atoms).

S	=	1(32.07 u)	=	32.07 u
O	=	3(16.00 u)	=	48.00 u
				80.07 u

b) Formula is $\mathbf{C_3H_8}$. Since it contains only carbon and hydrogen it is a hydrocarbon and with three carbons its name is **propane**.

C	=	3(12.01 u)	=	36.03 u
H	=	8(1.008 u)	=	8.064 u
				44.09 u

2.109 Plan: Use the chemical symbols and count the atoms of each type to give a molecular formula. Use the nomenclature rules in the chapter to derive the name. The molecular (formula) mass is the sum of the masses of each atom times its atomic mass.
Solution:
a) Formula is $\mathbf{N_2O}$. Name is **dinitrogen monoxide** (the prefix di- indicates 2 nitrogen atoms and mono- indicates 1 oxygen atom).

N	=	2(14.01 u)	=	28.02 u
O	=	1(16.00 u)	=	16.00 u
				44.02 u

b) Formula is $\mathbf{C_2H_6}$. Since it contains only carbon and hydrogen it is a hydrocarbon and with three carbons its name is **ethane**.

C	=	2(12.01 u)	=	24.02 u
H	=	6(1.008 u)	=	6.048 u
				30.07 u

2.110 Plan: Review the nomenclature rules in the chapter. For ionic compounds, name the metal, followed by the nonmetal name with an -ide suffix. For ionic compounds containing polyatomic ions, name the metal, followed by the name of the polyatomic ion. For metals, like many transition metals, that can form more than one ion each with a different charge, the ionic charge of the metal ion is indicated by a Roman numeral within parentheses immediately following the metal's name. Oxoacids (H + an oxoanion) are named by changing the suffix of the oxoanion: -ate becomes -ic acid and -ite becomes -ous acid. Greek numerical prefixes are used to indicate the number of atoms of each element in a compound composed of two nonmetals.
Solution:
a) blue vitriol $CuSO_4•5H_2O$ **copper(II) sulfate pentahydrate**
SO_4^{2-} = sulfate; II is used to indicate the 2+ charge of Cu; penta- is used to indicate the 5 waters of hydration.
b) slaked lime $Ca(OH)_2$ **calcium hydroxide**
The anion OH^- is hydroxide.
c) oil of vitriol H_2SO_4 **sulfuric acid**
SO_4^{2-} is the sulfate ion; since this is an acid, -ate becomes -ic acid.
d) washing soda Na_2CO_3 **sodium carbonate**
CO_3^{2-} is the carbonate ion.
e) muriatic acid HCl **hydrochloric acid**
Binary acids (H plus one other nonmetal) are named hydro- + nonmetal root + -ic acid.
f) Epsom salts $MgSO_4•7H_2O$ **magnesium sulfate heptahydrate**
SO_4^{2-} = sulfate; hepta- is used to indicate the 7 waters of hydration.
g) chalk $CaCO_3$ **calcium carbonate**
CO_3^{2-} is the carbonate ion.
h) dry ice CO_2 **carbon dioxide**
The prefix di- indicates 2 oxygen atoms; since there is only one carbon atom, no prefix is used.
i) baking soda $NaHCO_3$ **sodium hydrogen carbonate**
HCO_3^- is the hydrogen carbonate ion.
j) lye NaOH **sodium hydroxide**
The anion OH^- is hydroxide.

2.111 Plan: Use the chemical symbols and count the atoms of each type to give a molecular formula. Use the nomenclature rules in the chapter to derive the name. The molecular (formula) mass is the sum of the masses of each atom times its atomic mass.

Solution:

a) Each molecule has 2 blue spheres and 1 red sphere so the molecular formula is **N_2O**. This compound is composed of two nonmetals. The element with the lower group number is named first. Greek numerical prefixes are used to indicate the number of atoms of each element in the compound. The prefix di- indicates 2 nitrogen atoms and mono- indicates 1 oxygen atom. The name is **dinitrogen monoxide**.

N	=	2(14.01 u)	=	28.02u
O	=	1(16.00 u)	=	16.00 u
				44.02 u

b) Each molecule has 2 green spheres and 1 red sphere so the molecular formula is **Cl_2O**. This compound is composed of two nonmetals. When a compound contains oxygen and a halogen, the halogen is named first. Greek numerical prefixes are used to indicate the number of atoms of each element in the compound. The prefix di- indicates 2 chlorine atoms and mono- indicates 1 oxygen atom. The name is **dichlorine monoxide**.

Cl	=	2(35.45 u)	=	70.90 u
O	=	1(16.00 u)	=	16.00 u
				86.90 u

2.112 Plan: Review the discussion on separations.

Solution:

Separating the components of a mixture requires physical methods only; that is, no chemical changes (no changes in composition) take place and the components maintain their chemical identities and properties throughout. Separating the components of a compound requires a chemical change (change in composition).

2.113 Plan: Review the definitions of homogeneous and heterogeneous.

Solution:

A homogeneous mixture is uniform in its macroscopic, observable properties; a heterogeneous mixture shows obvious differences in properties (density, colour, state, etc.) from one part of the mixture to another.

2.114 A solution (such as salt or sugar dissolved in water) is a homogeneous mixture.

2.115 Plan: Review the definitions of homogeneous and heterogeneous. The key is that a homogeneous mixture has a uniform composition while a heterogeneous mixture does not. A mixture consists of two or more substances physically mixed together while a compound is a pure substance.

Solution:

a) Distilled water is a **compound** that consists of H_2O molecules only.
b) Gasoline is a **homogeneous mixture** of hydrocarbon compounds of uniform composition that can be separated by physical means (distillation).
c) Beach sand is a **heterogeneous mixture** of different size particles of minerals and broken bits of shells.
d) Wine is a **homogeneous mixture** of water, alcohol, and other compounds that can be separated by physical means (distillation).
e) Air is a **homogeneous mixture** of different gases, mainly N_2, O_2, and Ar.

2.116 Plan: Review the definitions of homogeneous and heterogeneous. The key is that a homogeneous mixture has a uniform composition while a heterogeneous mixture does not. A mixture consists of two or more substances physically mixed together while a compound is a pure substance.

Solution:

a) Orange juice is a **heterogeneous mixture** of water, juice, and bits of orange pulp.
b) Vegetable soup is a **heterogeneous mixture** of water, broth, and vegetables.
c) Cement is a **heterogeneous mixture** of various substances.
d) Calcium sulfate is a **compound** of calcium, sulfur, and oxygen in a fixed proportion.
e) Tea is a **homogeneous mixture**.

2.117 Plan: Review the discussion on separations.
Solution:
a) Filtration — separating the mixture on the basis of differences in particle size. The water moves through the holes in the colander but the larger pasta cannot.
b) **Extraction** — The coloured impurities are extracted into a solvent that is rinsed away from the raw sugar (or **chromatography**). A sugar solution is passed through a column in which the impurities stick to the stationary phase and the sugar moves through the column in the mobile phase.

2.118 Analysis time can be shortened by operating the column at a higher temperature or by increasing the rate of flow of the gaseous mobile phase.

2.119 Plan: Use the equation for the volume of a sphere in part a) to find the volume of the nucleus and the volume of the atom. Calculate the fraction of the atom volume that is occupied by the nucleus. For part b), calculate the total mass of the two electrons; subtract the electron mass from the mass of the atom to find the mass of the nucleus. Then calculate the fraction of the atom's mass contributed by the mass of the nucleus.
Solution:
a) Volume (m^3) of nucleus = $\frac{4}{3}\pi r^3 = \frac{4}{3}\pi\left(2.5\text{x}10^{-15}\text{ m}\right)^3 = 6.54498\text{x}10^{-44}\text{ m}^3$

Volume (m^3) of atom = $\frac{4}{3}\pi r^3 = \frac{4}{3}\pi\left(3.1\text{x}10^{-11}\text{ m}\right)^3 = 1.24788\text{x}10^{-31}\text{ m}^3$

Fraction of volume = $\frac{\text{volume of Nucleus}}{\text{volume of Atom}} = \frac{6.54498\text{x}10^{-44}\text{ m}^3}{1.24788\text{x}10^{-31}\text{ m}^3} = 5.2449\text{x}10^{-13} = \mathbf{5.2x10^{-13}}$

b) Mass of nucleus = mass of atom – mass of electrons
= $6.64648\text{x}10^{-24}\text{ g} - 2(9.10939\text{x}10^{-28}\text{ g}) = 6.64466\text{x}10^{-24}\text{ g}$

Fraction of mass = $\frac{\text{mass of Nucleus}}{\text{mass of Atom}} = \frac{\left(6.64466\text{x}10^{-24}\text{ g}\right)}{\left(6.64648\text{x}10^{-24}\text{ g}\right)} = 0.99972617 = \mathbf{0.999726}$

As expected, the volume of the nucleus relative to the volume of the atom is small while its relative mass is large.

2.120 Plan: Use Coulomb's law which states that the energy of attraction in an ionic bond is directly proportional to the *product of charges* and inversely proportional to the *distance between charges*. Choose the largest ionic charges and smallest radii for the strongest ionic bonding and the smallest ionic charges and largest radii for the weakest ionic bonding.
Solution:
Strongest ionic bonding: **MgO**. Mg^{2+}, Ba^{2+}, and O^{2-} have the largest charges. Attraction increases as distance decreases, so the positive ion with the smaller radius, Mg^{2+}, will form a stronger ionic bond than the larger ion Ba^{2+}.
Weakest ionic bonding: **RbI**. K^+, Rb^+, Cl^-, and I^- have the smallest charges. Attraction decreases as distance increases, so the ions with the larger radii, Rb^+ and I^-, will form the weakest ionic bond.

2.121 Plan: Use the chemical symbols and count the atoms of each type to give a molecular formula. Use the nomenclature rules in the chapter to derive the name. These compounds are composed of two nonmetals. Greek numerical prefixes are used to indicate the number of atoms of each element in each compound. The molecular (formula) mass is the sum of the masses of each atom times its atomic mass.
Solution:
a) Formula is BrF_3. When a compound is composed of two elements from the same group, the element with the higher period number is named first. The prefix tri- indicates 3 fluorine atoms. A prefix is used with the first word in the name only when more than one atom of that element is present. The name is **bromine trifluoride**.

Br	=	1(79.90 u)	=	79.90 u
F	=	3(19.00 u)	=	57.00 u
				136.90 u

b) The formula is SCl_2. The element with the lower group number is the first word in the name. The prefix di- indicates 2 chlorine atoms. A prefix is used with the first word in the name only when more than one atom of that element is present. The name is **sulfur dichloride**.

S	=	1(32.07 u)	=	32.07 u
Cl	=	2(35.45 u)	=	70.90 u
				102.97 u

c) The formula is PCl_3. The element with the lower group number is the first word in the name. The prefix tri- indicates 3 chlorine atoms. A prefix is used with the first word in the name only when more than one atom of that element is present. The name is **phosphorus trichloride**.

P	=	1(30.97 u)	=	30.97 u
Cl	=	3(35.45 u)	=	106.35 u
				137.32 u

d) The formula is N_2O_5. The element with the lower group number is the first word in the name. The prefix di- indicates 2 nitrogen atoms and the prefix penta- indicates 5 oxygen atoms. Only the second element is named with the suffix -ide. The name is **dinitrogen pentoxide**.

N	=	2(14.01 u)	=	28.02 u
O	=	5(16.00 u)	=	80.00 u
				108.02 u

2.122 Plan: These polyatomic ions are oxoanions composed of oxygen and another nonmetal. Oxoanions with the same number of oxygen atoms and nonmetals in the same group will have the same suffix ending. Only the nonmetal root name will change.

Solution:

a) SeO_4^{2-}	**selenate ion**	from SO_4^{2-} = sulfate ion
b) AsO_4^{3-}	**arsenate ion**	from PO_4^{3-} = phosphate ion
c) BrO_2^-	**bromite ion**	from ClO_2^- = chlorite ion
d) $HSeO_4^-$	**hydrogen selenate ion**	from HSO_4^- = hydrogen sulfate ion
e) TeO_3^{2-}	**tellurite ion**	from SO_3^{2-} = sulfite ion

2.123 Plan: Write the formula of the compound and find the molecular mass. Determine the mass percent of nitrogen or phosphorus by dividing the mass of nitrogen or phosphorus in the compound by the molecular mass and multiplying by 100. For part b), multiply the 100. g sample of compound by the mass ratio of ammonia to compound.

Solution:

a) Ammonium is NH_4^+ and dihydrogen phosphate is $H_2PO_4^-$. The formula is $NH_4H_2PO_4$.

N	=	1(14.01 u)	=	14.01 u
H	=	6(1.008 u)	=	6.048 u
P	=	1(30.97 u)	=	30.97 u
O	=	4(16.00 u)	=	64.00 u
				115.03 u

Mass percent of N = $\frac{14.01 \text{ u N}}{115.03 \text{ u compound}}(100)$ = **12.18% N**

Mass percent of P = $\frac{30.97 \text{ u P}}{115.03 \text{ u compound}}(100)$ = **26.92% P**

b) Mass (g) of ammonia (NH_3) = $(100.\text{ g } NH_4H_2PO_4)\left(\frac{17.03 \text{ u } NH_3}{115.03 \text{ u } NH_4H_2PO_4}\right)$ = **14.80 g NH_3**

2.124 Plan: Determine the percent oxygen in each oxide by subtracting the percent nitrogen from 100%. Express the percentage in u and divide by the atomic mass of the appropriate elements. Then divide each amount by the smaller number and convert to the simplest whole-number ratio. To find the mass of oxygen per 1.00 g of nitrogen, divide the mass percentage of oxygen by the mass percentage of nitrogen.

Solution:

a) I (100.00 – 46.69 N)% = 53.31% O

$$\left(\frac{46.69\text{ u N}}{14.01\text{ u N}}\right) = 3.3326\text{ N} \qquad \left(\frac{53.31\text{ u O}}{16.00\text{ u O}}\right) = 3.3319\text{ O}$$

$$\frac{3.3326\text{ N}}{3.3319} = 1.0002\text{ N} \qquad \frac{3.3319\text{ O}}{3.3319} = 1.0000\text{ O}$$

The simplest whole-number ratio is **1:1 N:O**.

II (100.00 – 36.85 N)% = 63.15% O

$$\left(\frac{36.85\text{ u N}}{14.01\text{ u N}}\right) = 2.6303\text{ N} \qquad \left(\frac{63.15\text{ u O}}{16.00\text{ u O}}\right) = 3.9469\text{ O}$$

$$\frac{2.6303\text{ N}}{2.6303} = 1.0000\text{ mol N} \qquad \frac{3.9469\text{ O}}{2.6303} = 1.5001\text{ O}$$

The simplest whole-number ratio is 1:1.5 N:O = **2:3 N:O**.

III (100.00 – 25.94 N)% = 74.06% O

$$\left(\frac{25.94\text{ u N}}{14.01\text{ u N}}\right) = 1.8515\text{ N} \qquad \left(\frac{74.06\text{ u O}}{16.00\text{ u O}}\right) = 4.6288\text{ O}$$

$$\frac{1.8515\text{ N}}{1.8515} = 1.0000\text{ N} \qquad \frac{4.6288\text{ O}}{1.8515} = 2.5000\text{ O}$$

The simplest whole-number ratio is 1:2.5 N:O = **2:5 N:O.**

b) I $\left(\frac{53.31\text{ u O}}{46.69\text{ u N}}\right) = 1.1418 = \mathbf{1.14\text{ g O}}$

II $\left(\frac{63.15\text{ u O}}{36.85\text{ u N}}\right) = 1.7137 = \mathbf{1.71\text{ g O}}$

III $\left(\frac{74.06\text{ u O}}{25.94\text{ u N}}\right) = 2.8550 = \mathbf{2.86\text{ g O}}$

2.125 Plan: Recall that density = mass/volume.

Solution:

The mass of an atom of Pb is several times that of one of Al. Thus, the density of Pb would be expected to be several times that of Al if approximately equal numbers of each atom were occupying the same volume.

2.126 Plan: Review the law of mass conservation and law of definite composition. For each experiment, compare the mass values before and after each reaction and examine the ratios of the mass of reacted sodium to the mass of reacted chlorine.

Solution:

In each case, the mass of the starting materials (reactants) equals the mass of the ending materials (products), so the law of mass conservation is observed.

Case 1: 39.34 g + 60.66 g = 100.00 g

Case 2: 39.34 g + 70.00 g = 100.00 g + 9.34 g

Case 3: 50.00 g + 50.00 g = 82.43 g + 17.57 g

Each reaction yields the product NaCl, not Na_2Cl or $NaCl_2$ or some other variation, so the law of definite composition is observed. In each case, the ratio of the mass of sodium to the mass of chlorine in the compound is the same.

Case 1: Mass Na/mass Cl_2 = 39.34 g/60.66 g = 0.6485

Case 2: Mass of reacted Cl_2 = initial mass – excess mass = 70.00 g – 9.34 g = 60.66 g Cl_2

Mass Na/mass Cl_2 = 39.34 g/60.66 g = 0.6485

Case 3: Mass of reacted Na = initial mass – excess mass = 50.00 g – 17.57 g = 32.43 g Na

Mass Na/mass Cl_2 = 32.43 g/50.00 g = 0.6486

2.127 Plan: Recall the definitions of solid, liquid, gas (from Chapter 1), element, compound, and homogeneous and heterogeneous mixtures.

Solution:

a) Gas is the phase of matter that fills its container. A mixture must contain at least two different substances. B, F, G, and I each contain only one gas. **D and E** each contain a mixture; E is a mixture of two different gases while D is a mixture of a gas and a liquid of a second substance.

b) An element is a substance that cannot be broken down into simpler substances. **A, C, G, and I** are elements.

c) The solid phase has a very high resistance to flow since it has a fixed shape. **A** shows a solid element.

d) A homogeneous mixture contains two or more substances and has only one phase. **E and H** are examples of this. E is a homogeneous mixture of two gases and H is a homogeneous mixture of two liquid substances.

e) A liquid conforms to the container shape and forms a surface. **C** shows one element in the liquid phase.

f) A diatomic particle is a molecule composed of two atoms. **B and G** contain diatomic molecules of gas.

g) A compound can be broken down into simpler substances. **B and F** show molecules of a compound in the gas phase.

h) The compound shown in **F** has molecules composed of two white atoms and one blue atom for a 2:1 atom ratio.

i) Mixtures can be separated into the individual components by physical means. **D, E, and H** are each a mixture of two different substances.

j) A heterogeneous mixture like **D** contains at least two different substances with a visible boundary between those substances.

k) Compounds obey the law of definite composition. **B and F** depict compounds.

2.128 Plan: To find the mass percent divide the mass of each substance in mg by the amount of seawater in mg and multiply by 100. The percent of an ion is the mass of that ion divided by the total mass of ions.

Solution:

a) Mass (mg) of seawater = $(1\text{ kg})\left(\frac{1000\text{ g}}{1\text{ kg}}\right)\left(\frac{1000\text{ mg}}{1\text{ g}}\right) = 1\text{x}10^6\text{ mg}$

$$\text{Mass \%} = \left(\frac{\text{mass of substance}}{\text{mass of seawater}}\right)(100\%)$$

$$\text{Mass \% Cl}^- = \left(\frac{18{,}980\text{ mg Cl}^-}{1\text{x}10^6\text{ mg seawater}}\right)(100\%) = \mathbf{1.898\%\ Cl^-}$$

$$\text{Mass \% Na}^+ = \left(\frac{10.560\text{ mg Na}^+}{1\text{x}10^6\text{ mg seawater}}\right)(100\%) = \mathbf{1.056\%\ Na^+}$$

$$\text{Mass \% SO}_4^{2-} = \left(\frac{2650\text{ mg SO}_4^{2-}}{1\text{x}10^6\text{ mg seawater}}\right)(100\%) = \mathbf{0.265\%\ SO_4^{2-}}$$

$$\text{Mass \% Mg}^{2+} = \left(\frac{1270\text{ mg Mg}^{2+}}{1\text{x}10^6\text{ mg seawater}}\right)(100\%) = \mathbf{0.127\%\ Mg^{2+}}$$

$$\text{Mass \% Ca}^{2+} = \left(\frac{400\text{ mg Ca}^{2+}}{1\text{x}10^6\text{ mg seawater}}\right)(100\%) = \mathbf{0.04\%\ Ca^{2+}}$$

$$\text{Mass \% K}^+ = \left(\frac{380\text{ mg K}^+}{1\text{x}10^6\text{ mg seawater}}\right)(100\%) = \mathbf{0.038\%\ K^+}$$

$$\text{Mass \% HCO}_3^- = \left(\frac{140\text{ mg HCO}_3^-}{1\text{x}10^6\text{ mg seawater}}\right)(100\%) = \mathbf{0.014\%\ HCO_3^-}$$

The mass percents do not add to 100% since the majority of seawater is H_2O.

b) Total mass of ions in 1 kg of seawater

$= 18{,}980\text{ mg} + 10{,}560\text{ mg} + 2650\text{ mg} + 1270\text{ mg} + 400\text{ mg} + 380\text{ mg} + 140\text{ mg} = 34{,}380\text{ mg}$

$$\%\ Na^+ = \left(\frac{10{,}560\text{ mg }Na^+}{34{,}380\text{ mg total ions}}\right)(100) = 30.71553 = \mathbf{30.72\%}$$

c) Alkaline earth metal ions are Mg^{2+} and Ca^{2+} (Group 2 ions).
Total mass % = 0.127% Mg^{2+} + 0.04% Ca^{2+} = 0.167%
Alkali metal ions are Na^+ and K^+ (Group 1 ions). Total mass % = 1.056% Na^+ + 0.038% K^+ = 1.094%

$$\frac{\text{Mass \% of alkali metal ions}}{\text{Mass \% of alkaline earth metal ions}} = \frac{1.094\%}{0.167\%} = 6.6$$

Total mass percent for alkali metal ions is **6.6 times greater** than the total mass percent for alkaline earth metal ions. Sodium ions (alkali metal ions) are dominant in seawater.
d) Anions are Cl^-, SO_4^{2-}, and HCO_3^-.
Total mass % = 1.898% Cl^- + 0.265% SO_4^{2-} + 0.014% HCO_3^- = 2.177% anions
Cations are Na^+, Mg^{2+}, Ca^{2+}, and K^+.
Total mass % = 1.056% Na^+ + 0.127% Mg^{2+} + 0.04% Ca^{2+} + 0.038% K^+ = 1.2610 = 1.26% cations
The mass fraction of **anions** is larger than the mass fraction of cations. Is the solution neutral since the mass of anions exceeds the mass of cations? Yes, although the mass is larger, the number of positive charges equals the number of negative charges.

2.129 Plan: Review the mass laws in the chapter.
Solution:
The **law of mass conservation** is illustrated in this change. The first flask has six oxygen atoms and six nitrogen atoms. The same number of each type of atom is found in both of the subsequent flasks. The mass of the substances did not change. The **law of definite composition** is also illustrated. During both temperature changes, the same compound, N_2O, was formed with the same composition.

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2.130 Plan: First, count each type of atom present to produce a molecular formula. The molecular (formula) mass is the sum of the atomic masses of all of the atoms. Divide the mass of each element in the compound by the molecular mass and multiply by 100 to obtain the mass percent of each element.
Solution:
The molecular formula of succinic acid is $\mathbf{C_4H_6O_4}$.

C	=	4(12.01 u)	=	48.04 u
H	=	6(1.008 u)	=	6.048 u
O	=	4 (16.00 u)	=	64.00 u
				118.09 u

$$\%\ C = \left(\frac{48.04\text{ u C}}{118.088\text{ u}}\right)100\% = 40.6815 = \mathbf{40.68\%\ C}$$

$$\%\ H = \left(\frac{6.048\text{ u H}}{118.088\text{ u}}\right)100\% = 5.1216 = \mathbf{5.122\%\ H}$$

$$\% \text{ O} = \left(\frac{64.00 \text{ u O}}{118.088 \text{ u}}\right)100\% = 54.1969 = \mathbf{54.20\%\ O}$$

Check: Total = (40.68 + 5.122 + 54.20)% = 100.00% The answer checks.

2.131 Plan: The toxic level of fluoride ion for a 70-kg person is 0.2 g. Convert this mass to mg and use the concentration of fluoride ion in drinking water to find the volume of water that contains the toxic amount. Convert the volume of the reservoir to liters and use the concentration of 1mg of fluoride ion per liter of water to find the mass of sodium fluoride required.

Solution:

A 70-kg person would have to consume 0.2 mg of F^- to reach the toxic level.

$$\text{Mass (mg) of fluoride for a toxic level} = \left(0.2 \text{ g F}^-\right)\left(\frac{1 \text{ mg F}^-}{0.001 \text{ g F}^-}\right) = 200 \text{ mg F}^-$$

$$\text{Volume (L) of water} = \left(200 \text{ mg}\right)\left(\frac{1 \text{ L water}}{1 \text{ mg F}^-}\right) = 200 \text{ L} = \mathbf{2x10^2\ L\ water}$$

$$\text{Volume (L) of reservoir} = 8.50 \text{x} 10^7 \text{ gal}\left(\frac{4 \text{ qt}}{1 \text{ gal}}\right)\left(\frac{1 \text{ L}}{1.057 \text{ qt}}\right) = 3.26651\text{x}10^8 \text{ L}$$

The molecular mass of NaF = 22.99 u Na + 19.00 u F = 41.99 u. There are 19.00 mg of F^- in every 41.99 mg of NaF.

$$\text{Mass (kg) of NaF} = \left(3.216651\text{x}10^8 \text{ L}\right)\left(\frac{1 \text{ mg F}^-}{1 \text{ L H}_2\text{O}}\right)\left(\frac{41.99 \text{ mg NaF}}{19.00 \text{ mg F}^-}\right)\left(\frac{10^{-3} \text{ g}}{1 \text{ mg}}\right)\left(\frac{1 \text{ kg NaF}}{10^3 \text{ g NaF}}\right)$$

$$= 710.88 \text{ kg} = \mathbf{711\ kg\ NaF}$$

2.132 Plan: Z = the atomic number of the element. A is the mass number. To find the percent abundance of each Sb isotope, let x equal the fractional abundance of one isotope and (1 – x) equal the fractional abundance of the second isotope since the sum of the fractional abundances must equal 1. Remember that atomic mass = (isotopic mass of the first isotope x fractional abundance) + (isotopic mass of the second isotope x fractional abundance).

Solution:

a) Antimony is element 51so Z = 51. Isotope of mass 120.904 u has a mass number of 121: $\mathbf{^{121}_{51}Sb}$

Isotope of mass 122.904 u has a mass number of 123: $\mathbf{^{123}_{51}Sb}$

b) Let x = fractional abundance of antimony-121. This makes the fractional abundance of antimony-123 = 1 – x

x(120.904 u) + (1 – x) (122.904 u) = 121.8 u

120.904 u(x) + 122.904 u – 122.904 u(x) = 121.8 u

2x = 1.104

x = 0.552 = **0.55 fraction of antimony-121**

1 – x = 1 – 0.552 = **0.45 fraction of antimony-123**

2.133 Plan: List all possible combinations of the isotopes. Determine the masses of each isotopic composition. The molecule consisting of the lower abundance isotopes (N-15 and O-18) is the least common, and the one containing only the more abundant isotopes (N-14 and O-16) will be the most common.

Solution:

a) Formula	Mass (u)	b)
$\mathbf{^{15}N_2{}^{18}O}$	2(15 u N) + 18 u O = **48**	**least common**
$\mathbf{^{15}N_2{}^{16}O}$	2(15 u N) + 16u O = **46**	
$\mathbf{^{14}N_2{}^{18}O}$	2(14 u N) + 18 u O = **46**	
$\mathbf{^{14}N_2{}^{16}O}$	2(14 u N) + 16 u O = **44**	**most common**
$\mathbf{^{15}N^{14}N^{18}O}$	1(15 u N) + 1(14 u N) + 18 u O = **47**	
$\mathbf{^{15}N^{14}N^{16}O}$	1(15 u N) + 1(14 u N) + 16 u O = **45**	

2.134 Plan: Review the information about the periodic table in the chapter.

Solution:

a) Nonmetals are located in the upper-right portion of the periodic table: **Black, red, green, and purple**

b) Metals are located in the large left portion of the periodic table: **Brown and blue**

c) Some nonmetals, such as oxygen, chlorine, and argon, are gases: **Red, green, and purple**

d) Most metals, such as sodium and barium are solids; carbon is a solid: **Brown, blue, and black**

e) Nonmetals form covalent compounds; most noble gases do not form compounds:
Black and red or black and green or red and green

f) Nonmetals form covalent compounds; most noble gases do not form compounds:
Black and red or black and green or red and green

g) Metals react with nonmetals to form ionic compounds. For a compound with a formula of MX, the ionic charges of the metals and nonmetal must be equal in magnitude like Na^+ and Cl^- or Ba^{2+} and O^{2-}: **Brown and green or blue and red**

h) Metals react with nonmetals to form ionic compounds. For a compound with a formula of MX, the ionic charges of the metals and nonmetal must be equal in magnitude like Na^+ and Cl^- or Ba^{2+} and O^{2-}: **Brown and green or blue and red**

i) Metals react with nonmetals to form ionic compounds. For a compound with a formula of M_2X, the ionic charge of the nonmetal must be twice as large as that of the metal like Na^+ and O^{2-} or Ba^{2+} and C^{4-}: **Brown and red or blue and black**

j) Metals react with nonmetals to form ionic compounds. For a compound with a formula of MX_2, the ionic charge of the metal must be twice as large as that of the nonmetal like Ba^{2+} and Cl^-: **Blue and green**

k) Most Group 18 elements are unreactive: **Purple**

l) Different compounds often exist between the same two nonmetal elements. Since oxygen exists as O^{2-} or O_2^{2-}, metals can sometimes form more than one compound with oxygen: **Black and red or red and green or black and green or brown and red or blue and red**

2.135 Plan: To find the formula mass of potassium fluoride, add the atomic masses of potassium and fluorine. Fluorine has only one naturally occurring isotope, so the mass of this isotope equals the atomic mass of fluorine. The atomic mass of potassium is the weighted average of the two isotopic masses: (isotopic mass of isotope 1 x fractional abundance) + (isotopic mass of isotope 2 x fractional abundance).

Solution:

Average atomic mass of K =
(isotopic mass of ^{39}K x fractional abundance) + (isotopic mass of ^{41}K x fractional abundance)

Average atomic mass of K = $(38.9637\text{ u})\left(\frac{93.258\%}{100\%}\right) + (40.9618\text{ u})\left(\frac{6.730\%}{100\%}\right)$ = 39.093 u

The formula for potassium fluoride is KF, so its molecular mass is (39.093 + 18.9984)u = **58.091 u**

2.136 Plan: List all possible combinations of the isotopes. BF_3 contains either ^{10}B or ^{11}B. Determine the masses of each isotopic composition and also the masses of each molecule missing one, two, or all three F atoms.

Solution:

$^{10}B^{19}F_3$ = 10 u B + 3(19 u F) = **67 u**
$^{10}B^{19}F_2$ = 10 u B + 2(19 u F) = **48 u**
$^{10}B^{19}F$ = 10 u B + 19 u F = **29 u**
^{10}B = 10 u B = **10. u**
$^{11}B^{19}F_3$ = 11 u B + 3(19 u F) = **68 u**
$^{11}B^{19}F_2$ = 11 u B + 2(19 u F) = **49 u**
$^{11}B^{19}F$ = 11 u B + 19 u F = **30 u**
^{11}B = 11 u B = **11 u**

2.137 Plan: One molecule of NO is released per atom of N in the medicine. Divide the total mass of NO released by the molecular mass of the medicine and multiply by 100 for mass percent.

Solution:

NO = (14.01 + 16.00) u = 30.01 u

Nitroglycerin:

$C_3H_5N_3O_9$ = 3(12.01 u C) + 5(1.008 u H) + 3(14.01 u N) + 9(16.00 u O) = 227.10 u

In $C_3H_5N_3O_9$ (molecular mass = 227.10 u), there are 3 atoms of N; since 1 molecule of NO is released per atom of N, this medicine would release 3 molecules of NO. The molecular mass of NO = 30.01 u.

Mass percent of NO = $\frac{\text{total mass of NO}}{\text{mass of compound}}(100\%) = \frac{3(30.01\text{ u})}{227.10\text{ u}}(100\%)$ = 39.6433% = **39.64%**

Isoamyl nitrate:
$C_5H_{11}NO_3$ = 5(12.01 u C) + 11(1.008 u H) + 1(14.01 u N) + 3(16.00 u O) = 133.15 u
In $(CH_3)_2CHCH_2CH_2ONO_2$ (molecular mass = 133.15 u), there is one atom of N; since 1 molecule of NO is released per atom of N, this medicine would release 1 molecule of NO.

Mass percent of NO = $\frac{\text{total mass of NO}}{\text{mass of compound}}(100\%) = \frac{1(30.01\text{ u})}{133.15\text{ u}}(100\%)$ = 22.5385 %= **22.54%**

2.138 Plan: First, count each type of atom present to produce a molecular formula. Determine the mass fraction of each element. Mass fraction = $\frac{\text{total mass of the element}}{\text{molecular mass of TNT}}$. The mass of TNT multiplied by the mass fraction of each element gives the mass of that element.
Solution:
The molecular formula for TNT is $C_7H_5O_6N_3$. The molecular mass of TNT is:

C	=	7(12.01 u)	=	84.07 u
H	=	5(1.008 u)	=	5.040 u
O	=	6(16.00 u)	=	96.00 u
N	=	3(14.01 u)	=	42.03 u
				227.14 u

The mass fraction of each element is:

C = $\frac{84.07\text{ u}}{227.14\text{ u}}$ = 0.3701 C H = $\frac{5.040\text{ u}}{227.14\text{ u}}$ = 0.02219 H

O = $\frac{96.00\text{ u}}{227.14\text{ u}}$ = 0.4226 O N = $\frac{42.03\text{ u}}{227.14\text{ u}}$ = 0.1850 N

Masses of each element in 1.00 kg of TNT = mass fraction of element x 1.00kg.
Mass (kg) C = 0.3701 x 1.00 kg = **0.370 kg C**
Mass (kg) H = 0.02219 x 1.00 kg = **0.0222 kg H**
Mass (kg) O = 0.4226 x 1.00 kg = **0.423 kg O**
Mass (kg) N = 0.1850 x 1.00 kg = **0.185 kg N**

2.139 Plan: The superscript is the mass number, the sum of the number of protons and neutrons. Consult the periodic table to get the atomic number (the number of protons). The mass number – the number of protons = the number of neutrons. Divide the number of neutrons by the number of protons to obtain the *N*/*Z* ratio. For atoms, the number of protons and electrons are equal.
Solution:

	neutrons (*N*)	protons (*Z*)	*N*/*Z*
a) $^{144}_{62}Sm$	144 – 62 = 82	62	82/62 = **1.3**
b) $^{56}_{26}Fe$	56 – 26 = 30	26	30/26 = **1.2**
c) $^{20}_{10}Ne$	20 – 10 = 10	10	10/10 = **1.0**
d) $^{107}_{47}Ag$	107 – 47 = 60	47	60/47 = **1.3**

b)	neutrons	protons	electrons
$^{238}_{92}U$	238 – 92 = **146**	**92**	**92**
$^{234}_{92}U$	234 – 92 = **142**	**92**	**92**
$^{214}_{82}Pb$	214 – 82 = **132**	**82**	**82**
$^{210}_{82}Pb$	210 – 82 = **128**	**82**	**82**
$^{206}_{82}Pb$	206 – 82 = **124**	**82**	**82**

2.140 Plan: Determine the mass percent of platinum by dividing the mass of Pt in the compound by the molecular mass of the compound and multiplying by 100. For part b), divide the total amount of money available by the cost of Pt per gram to find the mass of Pt that can be purchased. Use the mass percent of Pt to convert from mass of Pt to mass of compound.

Solution:

a) The molecular formula for platinol is $Pt(NH_3)_2Cl_2$. Its molecular mass is:

Pt	=	1(195.1 u)	=	195.1 u
N	=	2 (14.01 u)	=	28.02 u
H	=	6(1.008 u)	=	6.048 u
Cl	=	2(35.45 u)	=	70.90 u
				300.1 u

$$\text{Mass \% Pt} = \frac{\text{mass of Pt}}{\text{molecular mass of compound}}(100\%) = \frac{195.1\text{ u}}{300.1\text{ u}}(100\%) = 65.012\ \% = \mathbf{65.01\%\ Pt}$$

$$\text{b) Mass (g) of Pt} = \left(\$1.00 \times 10^6\right)\left(\frac{1\text{ g Pt}}{\$32}\right) = 31{,}250\text{ g Pt}$$

$$\text{Mass (g) of platinol} = \left(31{,}250\text{ g Pt}\right)\left(\frac{100\text{ g platinol}}{65.01\text{ g Pt}}\right) = 4.8070\text{x}10^4\text{ g} = \mathbf{4.8x10^4\ g\ platinol}$$

2.141 Plan: Obtain the information from the periodic table. The period number of an element is its row number while the group number is its column number.

Solution:

a) Building-block elements:

Name	Symbol	Atomic number	Atomic mass	Period number	Group number
Hydrogen	H	1	1.008 u	1	1
Carbon	C	6	12.01 u	2	14
Nitrogen	N	7	14.01 u	2	15
Oxygen	O	8	16.00 u	2	16
b) Macronutrients:					
Sodium	Na	11	22.99 u	3	1
Magnesium	Mg	12	24.31 u	3	2
Potassium	K	19	39.10 u	4	1
Calcium	Ca	20	40.08 u	4	2
Phosphorus	P	15	30.97 u	3	15
Sulfur	S	16	32.07 u	3	16
Chlorine	Cl	17	35.45 u	3	17

2.142 Plan: Review the definitions of pure substance, element, compound, homogeneous mixture, and heterogeneous mixture.

Solution:

Matter is divided into two categories: **pure substances and mixtures**.
Pure substances are divided into elements and **compounds**.
Mixtures are divided into solutions (homogeneous mixtures) and **heterogeneous mixtures**.

2.143 Plan: A change is physical when there has been a change in physical form but not a change in composition. In a chemical change, a substance is converted into a different substance.

Solution:

1) Initially, all the molecules are present in blue-blue or red-red pairs. After the change, there are no red-red pairs, and there are now red-blue pairs. Changing some of the pairs means there has been a **chemical change**.

2) There are two blue-blue pairs and four red-blue pairs both before and after the change, thus no chemical change occurred. The different types of molecules are separated into different boxes. This is a **physical change**.

3) The identity of the box contents has changed from pairs to individuals. This requires a **chemical change**.

4) The contents have changed from all pairs to all triplets. This is a change in the identity of the particles, thus, this is a **chemical change**.

5) There are four red-blue pairs both before and after, thus there has been no change in the identity of the individual units. There has been a **physical change**.