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CHAPTER 2
ATOMS, MOLECULES, AND IONS

Problem Categories

Biological: 2.79, 2.80.

Conceptual: 2.31, 2.32, 2.33, 2.34, 2.61, 2.65, 2.71, 2.72, 2.81, 2.82, 2.93, 2.103, 2.113.

Descriptive: 2.24, 2.25, 2.26, 2.51, 2.52, 2.66, 2.70, 2.76, 2.84, 2.85, 2.86, 2.87, 2.88, 2.90, 2.91, 2.92, 2.94, 2.98, 2.101, 2.114.

Environmental: 2.122.

Organic: 2.47, 2.48, 2.69, 2.105, 2.107, 2.108, 2.115.

Difficulty Level

Easy: 2.7, 2.8, 2.13, 2.14, 2.15, 2.16, 2.23, 2.31, 2.32, 2.33, 2.43, 2.44, 2.45, 2.46, 2.47, 2.48, 2.62, 2.91, 2.92, 2.99, 2.100.

Medium: 2.17, 2.18, 2.24. 2.26, 2.34, 2.35, 2.36, 2.49, 2.50, 2.57, 2.58, 2.59, 2.60, 2.61, 2.63, 2.64, 2.65, 2.66, 2.67, 2.68, 2.69, 2.70, 2.71, 2.72, 2.73, 2.74, 2.75, 2.76, 2.79, 2.80, 2.81, 2.82, 2.83, 2.84, 2.85, 2.86, 2.88, 2.89, 2.90, 2.93, 2.94, 2.95, 2.96, 2.98, 2.101, 2.102, 2.103, 2.110, 2.112, 2.114, 2.115.

Difficult: 2.25, 2.77, 2.78, 2.87, 2.97, 2.104, 2.105, 2.106, 2.107, 2.108, 2.109, 2.111, 2.113.

2.7 First, convert 1 cm to picometers.

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

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

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

$$r_{\text{atom}} = 10^{4} r_{\text{nucleus}} = 10^{4} \times 2.0 \text{ cm} \times \frac{1 \text{ m}}{100 \text{ cm}} \times \frac{1 \text{ mi}}{1609 \text{ m}} = \mathbf{0.12\ mi}$$

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

$$A = 26 + 28 = \mathbf{54}$$

2.14 **Strategy:** The 239 in Pu-239 is the mass number. The **mass number (*A*)** is the total number of neutrons and protons present in the nucleus of an atom of an element. You can look up the atomic number (number of protons) on the periodic table.

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Solution:

mass number = number of protons + number of neutrons

number of neutrons = mass number − number of protons = 239 − 94 = **145**

2.15

Isotope	$^{3}_{2}He$	$^{4}_{2}He$	$^{24}_{12}Mg$	$^{25}_{12}Mg$	$^{48}_{22}Ti$	$^{79}_{35}Br$	$^{195}_{78}Pt$
No. Protons	2	2	12	12	22	35	78
No. Neutrons	1	2	12	13	26	44	117

2.16

Isotope	$^{15}_{7}N$	$^{33}_{16}S$	$^{63}_{29}Cu$	$^{84}_{38}Sr$	$^{130}_{56}Ba$	$^{186}_{74}W$	$^{202}_{80}Hg$
No. Protons	7	16	29	38	56	74	80
No. Neutrons	8	17	34	46	74	112	122
No. Electrons	7	16	29	38	56	74	80

2.17 **(a)** $^{23}_{11}Na$ **(b)** $^{64}_{28}Ni$

2.18 The accepted way to denote the atomic number and mass number of an element X is as follows:

$$^{A}_{Z}X$$

where,

A = mass number

Z = atomic number

(a) $^{186}_{74}W$ **(b)** $^{201}_{80}Hg$

2.23 Helium and selenium are nonmetals whose name ends with *ium*. (Tellerium is a metalloid whose name ends in *ium*.)

2.24 **(a)** Metallic character increases as you progress down a group of the periodic table. For example, moving down Group 4A, the nonmetal carbon is at the top and the metal lead is at the bottom of the group.

(b) Metallic character decreases from the left side of the table (where the metals are located) to the right side of the table (where the nonmetals are located).

2.25 The following data were measured at 20°C.

(a) Li (0.53 g/cm^3) K (0.86 g/cm^3) H_2O (0.98 g/cm^3)

(b) Au (19.3 g/cm^3) Pt (21.4 g/cm^3) Hg (13.6 g/cm^3)

(c) Os (22.6 g/cm^3)

(d) Te (6.24 g/cm^3)

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

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

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

2.33 **Elements:** N_2, S_8, H_2
Compounds: NH_3, NO, CO, CO_2, SO_2

2.34 There are more than two correct answers for each part of the problem.

(a) H_2 and F_2 **(b)** HCl and CO **(c)** S_8 and P_4
(d) H_2O and $C_{12}H_{22}O_{11}$ (sucrose)

2.35

Ion	Na^+	Ca^{2+}	Al^{3+}	Fe^{2+}	I^-	F^-	S^{2-}	O^{2-}	N^{3-}
No. protons	11	20	13	26	53	9	16	8	7
No. electrons	10	18	10	24	54	10	18	10	10

2.36 The **atomic number (*Z*)** is the number of protons in the nucleus of each atom of an element. You can find this on a periodic table. The number of **electrons** in an *ion* is equal to the number of protons minus the charge on the ion.

number of electrons (ion) = number of protons − charge on the ion

Ion	K^+	Mg^{2+}	Fe^{3+}	Br^-	Mn^{2+}	C^{4-}	Cu^{2+}
No. protons	19	12	26	35	25	6	29
No. electrons	18	10	23	36	23	10	27

2.37 F^- and N^{3-} (10 electrons), Ar and P^{3-} (18 electrons), Fe^{3+} and V (23 electrons), Sn^{4+} and Ag^+ (46 electrons).

2.38 **(a)** $^{52}_{25}Mn$ **(b)** $^{22}_{10}Ne$ **(c)** $^{107}_{47}Ag$ **(d)** $^{127}_{53}I$ **(e)** $^{239}_{94}Pu$

2.45 **(a)** Sodium ion has a +1 charge and oxide has a −2 charge. The correct formula is **Na_2O**.
(b) The iron ion has a +2 charge and sulfide has a −2 charge. The correct formula is **FeS**.

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(c) The correct formula is $\mathbf{Co_2(SO_4)_3}$

(d) Barium ion has a +2 charge and fluoride has a −1 charge. The correct formula is $\mathbf{BaF_2}$.

2.46 **(a)** The copper ion has a +1 charge and bromide has a −1 charge. The correct formula is **CuBr**.

(b) The manganese ion has a +3 charge and oxide has a −2 charge. The correct formula is $\mathbf{Mn_2O_3}$.

(c) We have the Hg_2^{2+} ion and iodide (I^-). The correct formula is $\mathbf{Hg_2I_2}$.

(d) Magnesium ion has a +2 charge and phosphate has a −3 charge. The correct formula is $\mathbf{Mg_3(PO_4)_2}$.

2.47 **(a)** CN **(b)** CH **(c)** C_9H_{20} **(d)** P_2O_5 **(e)** BH_3

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

Solution:

(a) Dividing both subscripts by 2, the simplest whole number ratio of the atoms in Al_2Br_6 is $\mathbf{AlBr_3}$.

(b) Dividing all subscripts by 2, the simplest whole number ratio of the atoms in $Na_2S_2O_4$ is $\mathbf{NaSO_2}$.

(c) The molecular formula as written, $\mathbf{N_2O_5}$, contains the simplest whole number ratio of the atoms present. In this case, the molecular formula and the empirical formula are the same.

(d) The molecular formula as written, $\mathbf{K_2Cr_2O_7}$, contains the simplest whole number ratio of the atoms present. In this case, the molecular formula and the empirical formula are the same.

2.49 The molecular formula of glycine is $\mathbf{C_2H_5NO_2}$.

2.50 The molecular formula of ethanol is $\mathbf{C_2H_6O}$.

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

Ionic: LiF, $BaCl_2$, KCl

Molecular: $SiCl_4$, B_2H_6, C_2H_4

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

Ionic: NaBr, BaF_2, CsCl.

Molecular: CH_4, CCl_4, ICl, NF_3

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2.59

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

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

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

Solution:

(a) This is an ionic compound in which the metal cation (K^{+}) has only one charge. The correct name is **potassium hypochlorite**. Hypochlorite is a polyatomic ion with one less O atom than the chlorite ion, ClO_2^{-}.

(b) **silver carbonate**

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

(d) **potassium permanganate** **(e)** **cesium chlorate** **(f)** **hypoiodous acid**

(g) This is an ionic compound in which the metal can form more than one cation. Use a Roman numeral to specify the charge of the Fe ion. Since the oxide ion has a -2 charge, the Fe ion has a $+2$ charge. The correct name is **iron(II) oxide**.

(h) **iron(III) oxide**

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

(j) **sodium hydride** **(k)** **lithium nitride** **(l)** **sodium oxide**

(m) This is an ionic compound in which the metal cation (Na^{+}) has only one charge. The O_2^{2-} ion is called the peroxide ion. Each oxygen has a -1 charge. You can determine that each oxygen only has a -1

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charge, because each of the two Na ions has a +1 charge. Compare this to sodium oxide in part (l). The correct name is **sodium peroxide**.

(n) **iron(III) chloride hexahydrate**

2.61 **(a)** $RbNO_2$ **(b)** K_2S **(c)** NaHS **(d)** $Mg_3(PO_4)_2$ **(e)** $CaHPO_4$
(f) KH_2PO_4 **(g)** IF_7 **(h)** $(NH_4)_2SO_4$ **(i)** $AgClO_4$ **(j)** BCl_3

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

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

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

Solution:

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

(b) Strontium is an alkaline earth metal. It only forms a +2 cation. The polyatomic ion chlorite, ClO_2^-, has a −1 charge. Since the charges on the cation and anion are numerically different, the subscript of the cation is numerically equal to the charge on the anion, and the subscript of the anion is numerically equal to the charge on the cation. The correct formula is $\mathbf{Sr(ClO_2)_2}$.

(c) Perbromic tells you that the anion of this oxoacid is perbromate, BrO_4^-. The correct formula is $\mathbf{HBrO_4}(\boldsymbol{aq})$. Remember that (*aq*) means that the substance is dissolved in water.

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

(e) Na is an alkali metal. It only forms a +1 cation. The polyatomic ion ammonium, NH_4^+, has a +1 charge and the polyatomic ion phosphate, PO_4^{3-}, has a −3 charge. To balance the charge, you need 2 Na^+ cations. The correct formula is $\mathbf{Na_2(NH_4)PO_4}$.

(f) The Roman numeral II tells you that the Pb cation has a +2 charge. The polyatomic ion carbonate, CO_3^{2-}, has a −2 charge. Since, the charges are numerically equal, no subscripts are necessary in the formula. The correct formula is $\mathbf{PbCO_3}$.

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(g) The Roman numeral II tells you that the Sn cation has a +2 charge. Fluoride has a −1 charge. Since the charges on the cation and anion are numerically different, the subscript of the cation is numerically equal to the charge on the anion, and the subscript of the anion is numerically equal to the charge on the cation. The correct formula is $\mathbf{SnF_2}$.

(h) This is a molecular compound. The Greek prefixes tell you the number of each type of atom in the molecule. The correct formula is $\mathbf{P_4S_{10}}$.

(i) The Roman numeral II tells you that the Hg cation has a +2 charge. Oxide has a −2 charge. Since, the charges are numerically equal, no subscripts are necessary in the formula. The correct formula is **HgO**.

(j) The Roman numeral I tells you that the Hg cation has a +1 charge. However, this cation exists as Hg_2^{2+}. Iodide has a −1 charge. You need two iodide ions to balance the +2 charge of Hg_2^{2+}. The correct formula is $\mathbf{Hg_2I_2}$.

(k) This is a molecular compound. The Greek prefixes tell you the number of each type of atom in the molecule. The correct formula is $\mathbf{SeF_6}$.

2.63 Let's compare the ratio of the fluorine masses in the two compounds.

$$\frac{3.55 \text{ g F}}{2.37 \text{ g F}} = 1.50$$

This calculation indicates that there is 1.5 times more fluorine by mass in SF_6 compared to the other compound. The value of *n* is:

$$\boldsymbol{n} = \frac{6}{1.5} = \mathbf{4}$$

This is consistent with the Law of Multiple Proportions which states that if two elements combine to form more than one compound, the masses of one element that combine with a fixed mass of the other element are in ratios of small whole numbers. In this case, the ratio of the masses of fluorine in the two compounds is 6:4 or 3:2.

2.64 **(a)** dinitrogen pentoxide (N_2O_5)

(b) boron trifluoride (BF_3)

(c) dialuminum hexabromide (Al_2Br_6)

2.65 Uranium is radioactive. It loses mass because it constantly emits alpha (α) particles.

2.66 Changing the electrical charge of an atom usually has a major effect on its chemical properties. The two electrically neutral carbon isotopes should have nearly identical chemical properties.

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

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2.68 Atomic number = 127 − 74 = 53. This anion has 53 protons, so it is an iodide ion. Since there is one more electron than protons, the ion has a −1 charge. The correct symbol is $\mathbf{I^-}$.

2.69

(a)	molecular, C_3H_8 empirical, C_3H_8	**(b)**	molecular, C_2H_2 empirical, CH	**(c)**	molecular, C_2H_6 empirical, CH_3	**(d)**	molecular, C_6H_6 empirical, CH

2.70 NaCl is an ionic compound; it doesn't form molecules.

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

2.72 The species and their identification are as follows:

(a)	SO_2	molecule and compound	**(g)**	O_3	element and molecule
(b)	S_8	element and molecule	**(h)**	CH_4	molecule and compound
(c)	Cs	element	**(i)**	KBr	compound
(d)	N_2O_5	molecule and compound	**(j)**	S	element
(e)	O	element	**(k)**	P_4	element and molecule
(f)	O_2	element and molecule	**(l)**	LiF	compound

2.73

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

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

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

(d) **A**: $^{10}_{5}B$ **B**: $^{14}_{7}N^{3-}$ **C**: $^{39}_{19}K^{+}$ **D**: $^{66}_{30}Zn^{2+}$ **E**: $^{81}_{35}Br^{-}$ **F**: $^{11}_{5}B$ **G**: $^{19}_{9}F$

2.74

(a)	Ne, 10 p, 10 n	**(b)**	Cu, 29 p, 34 n	**(c)**	Ag, 47 p, 60 n
(d)	W, 74 p, 108 n	**(e)**	Po, 84 p, 119 n	**(f)**	Pu, 94 p, 140 n

2.75

(a)	BaO, barium oxide	**(b)**	Ca_3P_2, calcium phosphide
(c)	Al_2S_3, aluminum sulfide	**(d)**	Li_3N, lithium nitride

2.76 **(a)** Cu **(b)** P **(c)** Kr **(d)** Cs **(e)** Al **(f)** Sb **(g)** Cl **(h)** Sr

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2.77 When an anion is formed from an atom, you have the same number of protons attracting more electrons. The electrostatic attraction is weaker, which allows the electrons on average to move farther from the nucleus. An anion is larger than the atom from which it is derived. When a cation is formed from an atom, you have the same number of protons attracting fewer electrons. The electrostatic attraction is stronger, meaning that on average, the electrons are pulled closer to the nucleus. A cation is smaller than the atom from which it is derived.

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

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

$$V = \frac{4}{3}\pi r^3 = \frac{4}{3}\pi(3.04 \times 10^{-13}\ \text{cm})^3 = 1.177 \times 10^{-37}\ \text{cm}^3$$

The density of the nucleus can now be calculated.

$$d = \frac{\text{m}}{\text{V}} = \frac{3.82 \times 10^{-23}\ \text{g}}{1.177 \times 10^{-37}\ \text{cm}^3} = \mathbf{3.25 \times 10^{14}\ g/cm^3}$$

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

The mass of 11 electrons is:

$$11\ \text{electrons} \times \frac{9.1095 \times 10^{-28}\ \text{g}}{1\ \text{electron}} = 1.00205 \times 10^{-26}\ \text{g}$$

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

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

$$V_{\text{atom}} = \frac{4}{3}\pi r^3 = \frac{4}{3}\pi(1.86 \times 10^{-8}\ \text{cm})^3 = 2.695 \times 10^{-23}\ \text{cm}^3$$

$$V_{\text{electrons}} = V_{\text{atom}} - V_{\text{nucleus}} = (2.695 \times 10^{-23}\ \text{cm}^3) - (1.177 \times 10^{-37}\ \text{cm}^3) = 2.695 \times 10^{-23}\ \text{cm}^3$$

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

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

$$d = \frac{m}{V} = \frac{1.00205 \times 10^{-26}\,\text{g}}{2.695 \times 10^{-23}\,\text{cm}^3} = \mathbf{3.72 \times 10^{-4}\,g/cm^3}$$

The above results do support Rutherford's model. Comparing the space occupied by the electrons to the volume of the nucleus, it is clear that most of the atom is empty space. Rutherford also proposed that the nucleus was a *dense* central core with most of the mass of the atom concentrated in it. Comparing the density of the nucleus with the density of the space occupied by the electrons also supports Rutherford's model.

2.79 The molecular formula of caffeine is $\mathbf{C_8H_{10}N_4O_2}$. The empirical formula is $\mathbf{C_4H_5N_2O}$.

2.80 The empirical and molecular formulas of acetaminophen are $\mathbf{C_8H_9NO_2}$.

2.81 **(a)** Iodate ion is IO_3^{2-}. The correct formula is $\mathbf{Mg(IO_3)_2}$.

(b) The formula shown is phosphorous acid. The correct formula for phosphoric acid is $\mathbf{H_3PO_4}$.

(c) Sulfite ion is SO_3^{2-}. The correct formula is $\mathbf{BaSO_3}$.

(d) NH_4^+ is the ammonium ion. The correct formula is $\mathbf{NH_4HCO_3}$.

2.82 **(a)** The charge on the tin cation needs to be specified. The correct name is **tin(IV) chloride**.

(b) The charge on the copper ion is +1. The correct name is **copper(I) oxide**.

(c) The charge on the cobalt cation needs to be specified. The correct name is **cobalt(II) nitrate**.

(d) $Cr_2O_7^{2-}$ is the dichromate ion. The correct name is **sodium dichromate**.

2.83

Symbol	$^{11}_{5}B$	$^{54}_{26}Fe^{2+}$	$^{31}_{15}P^{3-}$	$^{196}_{79}Au$	$^{222}_{86}Rn$
Protons	5	26	15	79	86
Neutrons	6	28	16	117	136
Electrons	5	24	18	79	86
Net Charge	0	+2	−3	0	0

2.84 **(a)** Ionic compounds are typically formed between metallic and nonmetallic elements.

(b) In general the transition metals, the actinides, and the lanthanides have variable charges.

2.85 **(a)** Li^+, alkali metals always have a +1 charge in ionic compounds

(b) S^{2-}

(c) I^-, halogens have a −1 charge in ionic compounds

(d) N^{3-}

(e) Al^{3+}, aluminum always has a +3 charge in ionic compounds

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(f) Cs^+, alkali metals always have a +1 charge in ionic compounds

(g) Mg^{2+}, alkaline earth metals always have a +2 charge in ionic compounds.

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

2.87 The binary Group 7A element acids are: HF, hydrofluoric acid; HCl, hydrochloric acid; HBr, hydrobromic acid; HI, hydroiodic acid. Oxoacids containing Group 7A elements (using the specific examples for chlorine) are: $HClO_4$, perchloric acid; $HClO_3$, chloric acid; $HClO_2$, chlorous acid: HClO, hypochlorous acid.

Examples of oxoacids containing other Group A-block elements are: H_3BO_3, boric acid (Group 3A); H_2CO_3, carbonic acid (Group 4A); HNO_3, nitric acid and H_3PO_4, phosphoric acid (Group 5A); and H_2SO_4, sulfuric acid (Group 6A). Hydrosulfuric acid, H_2S, is an example of a binary Group 6A acid while HCN, hydrocyanic acid, contains both a Group 4A and 5A element.

2.88 Mercury (Hg) and bromine (Br_2)

2.89

(a)	Isotope	^4_2He	$^{20}_{10}Ne$	$^{40}_{18}Ar$	$^{84}_{36}Kr$	$^{132}_{54}Xe$
	No. Protons	2	10	18	36	54
	No. Neutrons	2	10	22	48	78
(b)	neutron/proton ratio	1.00	1.00	1.22	1.33	1.44

The neutron/proton ratio increases with increasing atomic number.

2.90 H_2, N_2, O_2, F_2, Cl_2, He, Ne, Ar, Kr, Xe, Rn

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

2.92 They do not have a strong tendency to form compounds. Helium, neon, and argon are chemically inert.

2.93 Magnesium and strontium are also alkaline earth metals. You should expect the charge of the metal to be the same (+2). **MgO** and **SrO**.

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

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2.95 **(a)** Berkelium (Berkeley, CA); Europium (Europe); Francium (France); Scandium (Scandinavia); Ytterbium (Ytterby, Sweden); Yttrium (Ytterby, Sweden).

(b) Einsteinium (Albert Einstein); Fermium (Enrico Fermi); Curium (Marie and Pierre Curie); Mendelevium (Dmitri Mendeleev); Lawrencium (Ernest Lawrence).

(c) Arsenic, Cesium, Chlorine, Chromium, Iodine.

2.96 The atomic number is 77 – 43 = 34. The symbol for the anion is $^{77}\mathbf{Se}^{2-}$.

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

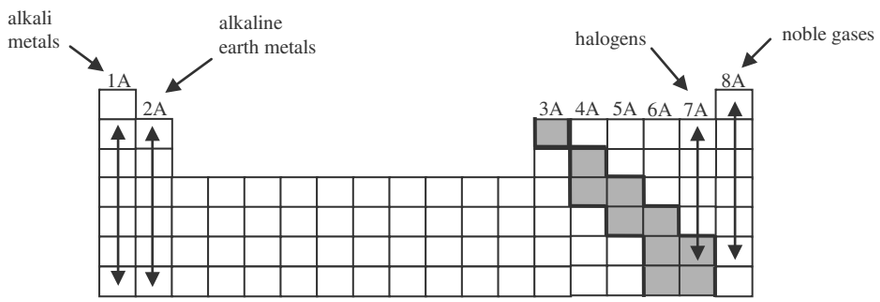
2.98 **(a)** NaH, sodium hydride **(b)** B_2O_3, diboron trioxide **(c)** Na_2S, sodium sulfide

(d) AlF_3, aluminum fluoride **(e)** OF_2, oxygen difluoride **(f)** $SrCl_2$, strontium chloride

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

2.100 All of these are molecular compounds. We use prefixes to express the number of each atom in the molecule. The names are nitrogen trifluoride (NF_3), phosphorus pentabromide (PBr_5), and sulfur dichloride (SCl_2).

2.101



The metalloids are shown in gray.

2.102

Cation	Anion	Formula	Name
$\mathbf{Mg^{2+}}$	$\mathbf{HCO_3^-}$	$\mathbf{Mg(HCO_3)_2}$	Magnesium bicarbonate
$\mathbf{Sr^{2+}}$	$\mathbf{Cl^-}$	$SrCl_2$	**Strontium chloride**
Fe^{3+}	NO_2^-	$\mathbf{Fe(NO_2)_3}$	**Iron(III) nitrite**

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Mn^{2+}	**ClO_3^-**	**$Mn(ClO_3)_2$**	Manganese(II) chlorate
Sn^{4+}	**Br^-**	$SnBr_4$	**Tin(IV) bromide**
Co^{2+}	PO_4^{3-}	**$Co_3(PO_4)_2$**	**Cobalt(II) phosphate**
Hg_2^{2+}	I^-	**Hg_2I_2**	**Mercury(I) iodide**
Cu^+	**CO_3^{2}**	Cu_2CO_3	**Copper(I) carbonate**
Li^+	**N^{3-}**	**Li_3N**	Lithium nitride
Al^{3+}	S^{2-}	**Al_2S_3**	**Aluminum sulfide**

2.103 **(a)** $CO_2(s)$, solid carbon dioxide
(b) NaCl, sodium chloride
(c) N_2O, nitrous oxide
(d) $CaCO_3$, calcium carbonate
(e) CaO, calcium oxide
(f) $Ca(OH)_2$, calcium hydroxide
(g) $NaHCO_3$, sodium bicarbonate
(h) $Na_2CO_3 \cdot 10H_2O$, sodium carbonate decahydrate
(i) $CaSO_4 \cdot 2H_2O$, calcium sulfate dihydrate
(j) $Mg(OH)_2$, magnesium hydroxide

2.104 The change in energy is equal to the energy released. We call this ΔE. Similarly, Δm is the change in mass. Because $m = \dfrac{E}{c^2}$, we have

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

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

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

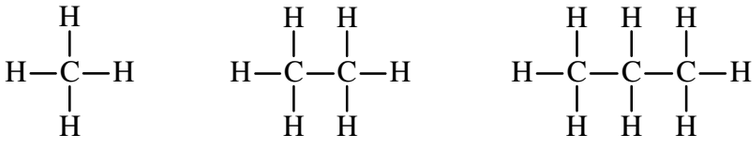
$$12.096 \text{ g} + 96.000 = 108.096 \text{ g}$$

The predicted change (loss) in mass is only 1.91×10^{-8} g which is too small a quantity to measure. Therefore, for all practical purposes, the law of conservation of mass is assumed to hold for ordinary chemical processes.

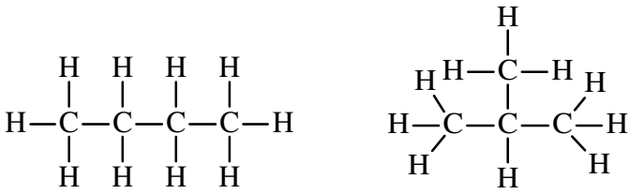
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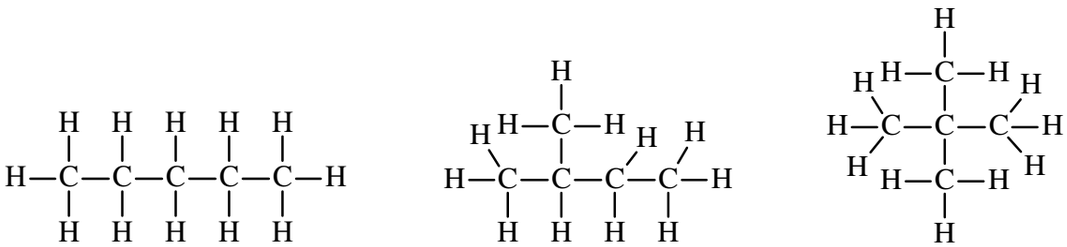
2.105 CH_4, C_2H_6, and C_3H_8 each only have one structural formula.



C_4H_{10} has two structural formulas.



C_5H_{12} has three structural formulas.



2.106 **(a)** The volume of a sphere is

$$V = \frac{4}{3}\pi r^3$$

Volume is proportional to the number of nucleons. Therefore,

$$V \propto \text{A (mass number)}$$

$$r^3 \propto \text{A}$$

$$\boldsymbol{r \propto A^{1/3}}$$

(b) Using the equation given in the problem, we can first solve for the radius of the lithium nucleus and then solve for its volume.

$$r = r_0 A^{1/3}$$

$$r = (1.2 \times 10^{-15}\ \text{m})(7)^{1/3}$$

$$r = 2.3 \times 10^{-15}\ \text{m}$$

$$V = \frac{4}{3}\pi r^3$$

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$$V_{\text{nucleus}} = \frac{4}{3}\pi(2.3 \times 10^{-15}\ \text{m})^3 = \mathbf{5.1 \times 10^{-44}\ m^3}$$

(c) In part (b), the volume of the nucleus was calculated. Using the radius of a Li atom, the volume of a Li atom can be calculated.

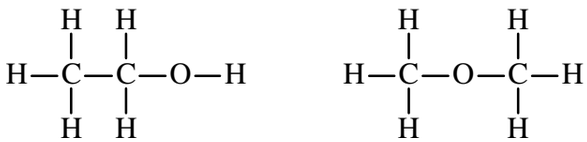
$$V_{\text{atom}} = \frac{4}{3}\pi r^3 = \frac{4}{3}\pi(152 \times 10^{-12}\ \text{m})^3 = 1.47 \times 10^{-29}\ \text{m}^3$$

The fraction of the atom's volume occupied by the nucleus is:

$$\frac{V_{\text{nucleus}}}{V_{\text{atom}}} = \frac{5.1 \times 10^{-44}\ \text{m}^3}{1.47 \times 10^{-29}\ \text{m}^3} = \mathbf{3.5 \times 10^{-15}}$$

Yes, this calculation shows that the volume of the nucleus is much, much smaller than the volume of the atom, which supports Rutherford's model of an atom.

2.107 Two different structural formulas for the molecular formula C_2H_6O are:



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

2.108 **(a)**

Ethane	Acetylene
2.65 g C	4.56 g C
0.665 g H	0.383 g H

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

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

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

$$\frac{1.14\ \text{g}}{0.383\ \text{g}} \approx 3$$

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

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

Ethane	Acetylene
CH_3	CH
C_2H_6	C_2H_2

2.109 **(a)** The following strategy can be used to convert from the volume of the Pt cube to the number of Pt atoms.

$$\text{cm}^3 \rightarrow \text{grams} \rightarrow \text{atoms}$$

$$1.0 \text{ cm}^3 \times \frac{21.45 \text{ g Pt}}{1 \text{ cm}^3} \times \frac{1 \text{ atom Pt}}{3.240 \times 10^{-22} \text{ g Pt}} = \mathbf{6.6 \times 10^{22} \text{ Pt atoms}}$$

(b) Since 74 percent of the available space is taken up by Pt atoms, 6.6×10^{22} atoms occupy the following volume:

$$0.74 \times 1.0 \text{ cm}^3 = 0.74 \text{ cm}^3$$

We are trying to calculate the radius of a single Pt atom, so we need the volume occupied by a single Pt atom.

$$\text{volume Pt atom} = \frac{0.74 \text{ cm}^3}{6.6 \times 10^{22} \text{ Pt atoms}} = 1.12 \times 10^{-23} \text{ cm}^3/\text{Pt atom}$$

The volume of a sphere is $\frac{4}{3}\pi r^3$. Solving for the radius:

$$V = 1.12 \times 10^{-23} \text{ cm}^3 = \frac{4}{3}\pi r^3$$

$$r^3 = 2.67 \times 10^{-24} \text{ cm}^3$$

$$r = 1.4 \times 10^{-8} \text{ cm}$$

Converting to picometers:

$$\textbf{radius Pt atom} = (1.4 \times 10^{-8} \text{ cm}) \times \frac{0.01 \text{ m}}{1 \text{ cm}} \times \frac{1 \text{ pm}}{1 \times 10^{-12} \text{ m}} = \mathbf{1.4 \times 10^2 \text{ pm}}$$

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

$$\text{Mass number} = \text{number of protons} + \text{number of neutrons}$$

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

$$55 = A + 1.2A$$
$$A = 25$$

The element with atomic number 25 is **manganese**, **Mn**.

2.111

S	N
B	I

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

2.113 From the equation derived in Problem 2.106, we can first solve for the radius of the iron nucleus and then solve for its volume.

$$r = r_0 A^{1/3}$$

$$r = (1.2 \times 10^{-15}\ \text{m})(56)^{1/3}$$

$$r = 4.6 \times 10^{-15}\ \text{m}$$

$$V = \frac{4}{3}\pi r^3$$

$$V_{\text{nucleus}} = \frac{4}{3}\pi(4.6 \times 10^{-15}\ \text{m})^3 = 4.1 \times 10^{-43}\ \text{m}^3 = 4.1 \times 10^{-37}\ \text{cm}^3$$

The density of an iron-56 nucleus is:

$$\textbf{density} = \frac{\text{mass}}{\text{volume}} = \frac{9.229 \times 10^{-23}\ \text{g}}{4.1 \times 10^{-37}\ \text{cm}^3} = \mathbf{2.3 \times 10^{14}\ g/cm^3}$$

The density of an iron nucleus is 230 trillion grams per 1 cm^3. Because of this very high density, neutrons are needed to stabilize the nucleus, keeping the protons from repelling each other. It has been postulated that the

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exterior of the neutron is negatively charged and the interior is positively charged. The negative exteriors of the neutrons attract the protons holding the particles together in this extremely dense arrangement.

2.114 The formula of the ionic compound is $\mathbf{XY_2}$. Element X is most likely in Group 4B and element Y is most likely in Group 6A. A possible compound is TiO_2, **titanium(IV) oxide**. Other choices are elements in Group 4A: SnO_2 [tin(IV) oxide] and PbO_2 [lead(IV) oxide].

2.115 Let's compare the ratio of the hydrogen masses in the three compounds. To do this, we need to start with the same mass of carbon. If we were to start with 0.749 g of C in ethane, how much hydrogen would combine with 0.749 g of carbon? If we were to start with 0.749 g of C in propane, how much hydrogen would combine with 0.749 g of carbon?

$$\text{ethane}: \quad 0.201 \text{ g H} \times \frac{0.749 \text{ g C}}{0.799 \text{ g C}} = 0.188 \text{ g H}$$

$$\text{propane}: \quad 0.183 \text{ g H} \times \frac{0.749 \text{ g C}}{0.817 \text{ g C}} = 0.168 \text{ g H}$$

We can calculate the ratio of H between methane and propane, and ethane and propane.

$$\text{methane : propane}: \quad \frac{0.251 \text{ g}}{0.168 \text{ g}} = 1.49$$

$$\text{ethane : propane}: \quad \frac{0.188 \text{ g}}{0.168 \text{ g}} = 1.12$$

This is consistent with the Law of Multiple Proportions which states that if two elements combine to form more than one compound, the masses of one element (in this case hydrogen) that combine with a fixed mass of the other element (in this case carbon) are in ratios of small whole numbers. The ratio of the masses of hydrogen in the three compounds, propane, ethane, and methane is 1:1.12:1.49 or **8:9:12**.

2.116 **Basic approach:**

- Determine the charge of an alpha particle (helium nucleus) and gold nucleus in coulombs.
- Use Coulomb's law to solve for the distance at which the electrical potential energy is equal to the initial kinetic energy.

When the α particle comes to a stop, its kinetic energy is converted to potential energy. The charge of a proton is $+1.6022 \times 10^{-19}$ C. An α particle which contains 2 protons and 2 neutrons has a charge, Q_1, of 3.2044×10^{-19} C. A gold nucleus which contains 79 protons has a charge, Q_2, of 1.2657×10^{-17} C. We solve for r, the distance of separation in meters between the alpha particle and the gold nucleus.

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$$r = \frac{kQ_1Q_2}{E} = \frac{\left(9.0 \times 10^9 \dfrac{\cancel{\text{kg}} \cdot \text{m}^{\cancel{3}}}{\cancel{s^2} \cdot \cancel{C^2}}\right)(3.2044 \times 10^{-19}\ \cancel{C})(1.2657 \times 10^{-17}\ \cancel{C})}{\left(6.0 \times 10^{-14} \dfrac{\text{kg} \cdot \text{m}^{\cancel{2}}}{\cancel{s^2}}\right)} = \mathbf{6.1 \times 10^{-13}\ m}$$

2.117 **Basic approach:**

- Consider that almost all of the volume in an atom or monatomic ion is occupied by the electrons.
- Determine the number of electrons in each species, and rank their size according the number of electrons.

The volume of each species depends on the outer boundary of the electrons. The number of protons remains the same in all three cases. Li^- has the most electrons and therefore the greatest repulsion among them, leading to a larger outer boundary compared to the Li atom, which has one fewer electron. Li^+ has the fewest number of electrons and the smallest outer boundary. Thus, the size decreases as follows:

$$Li^- > Li > Li^+$$

In general for a given element, the anion is the largest, the atom is second, and the cation is the smallest. In Chapter 8 we will consider a more sophisticated way of predicting the relative sizes of atoms and monatomic ions based on electron configurations and nuclear charge.

2.118 Isotopes of the same element have the same number of electrons, hence following the reasoning in Problem 2.117, they would have the same size. The number of neutrons affects the size of the nucleus, but not the size of the atom.

2.119 **Basic approach:**

- Look up the size of a silver atom.
- Calculate the number of silver atoms required to reach the minimum length visible to the human eye, and then calculate the number of silver atoms required to give a surface area that would be visible.

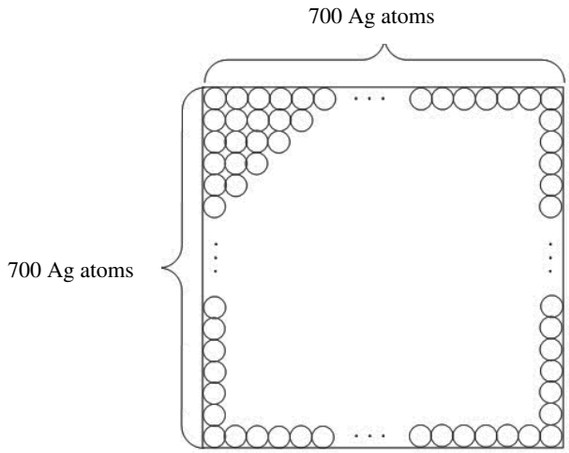
A Web search shows that the atomic radius of an Ag atom is 144 pm or 1.44×10^{-8} cm. The number of Ag atoms that must be lined up along one dimension is given by the minimum length humans can see divided by the diameter of a silver atom

$$\frac{2 \times 10^{-5}\ \text{cm}}{2 \times 1.44 \times 10^{-8}\ \text{cm/Ag atom}} \approx 700\ \text{Ag atoms}$$

But to actually be seen by the human eye, a grouping of silver atoms would need to have both length and width. Assuming a square array of 700 Ag atoms by 700 Ag atoms the number of silver atoms required to be visible would be $700 \times 700 \approx 500{,}000$ Ag atoms.

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2.120 Basic approach:

• Estimate the radius of a pea.

• Look up the radius of a typical atom and its nucleus. The radius of the atom provides a range of possible distances of the electron from the nucleus.

• Use the ratio of the radii for the pea and the atom to calculate the distance of the electron from the nucleus if the nucleus were the size of a pea.

The diameter of a pea is about 0.5 cm. Radius of nucleus is about 5×10^{-13} cm. Therefore, the expansion factor for the radius of the nucleus is

$$\frac{0.5 \text{ cm}/2}{5 \times 10^{-13} \text{ cm}} = 5 \times 10^{11} \text{ times}$$

A typical atom has a radius of about 100 pm (10^{-8} cm). Using that value as the distance of the electron from the nucleus in an atom, then if the atom was scaled to the size of a pea, the electron would be $(1 \times 10^{-8} \text{ cm}) \times (5 \times 10^{11}) = 5 \times 10^{3}$ cm or 50 m (roughly half a football field) from the nucleus!

2.121 Basic approach:

• Consider the relative composition of sodium and potassium ions in biological systems (for starters see the Chemistry in Action on p. 49, and then look up the relative compositions in other biological systems).

Over long periods of time (on a geological scale), minerals containing potassium and sodium are slowly decomposed by wind and rain, and their K^+ and Na^+ ions are converted to more soluble compounds. Eventually, rain leaches these compounds out of the soil and carries them to the sea. Plants take up many of the K^+ ions along the way, while the Na^+ ions are free to move on to the sea because they are not needed for biological functions by the plants.

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2.122 Basic approach:

- Estimate the diameter of a typical carbon microsphere and calculate the volume.
- Estimate for the density of carbon in the microsphere.
- Look up the most common isotope of carbon and calculate the mass of that isotope.
- Calculate the number of atoms in the microsphere.

A typical microsphere appears to be about 3 μm in diameter, giving a volume of

$$\frac{4}{3}\pi\left(\frac{d}{2}\right)^3 = \frac{4}{3}\pi\left(\frac{3\ \mu\text{m}}{2}\right)^3 \times \left(\frac{10^2\ \text{cm}}{10^6\ \mu\text{m}}\right)^3 \approx 1\times10^{-11}\ \text{cm}^3$$

It is reasonable to expect that the density of carbon in the microspheres would be somewhere between that of graphite (2.2 g/cm^3) and diamond (3.5 g/cm^3), so we assume a density of 3 g/cm^3 for the carbon in the microsphere. The mass of a typical carbon microsphere would be

$$1\times10^{-11}\ \text{cm}^3 \times \frac{3\ \text{g C}}{\text{cm}^3} = 3\times10^{-11}\ \text{g C}$$

While this may seem like an exceedingly small mass, it still represents a very large number of carbon atoms which are much smaller and far less massive than the microspheres. In order to determine the number of atoms in a typical microsphere, we need to approximate the mass of an individual carbon atom. An online search tells us that the most common isotope of carbon is $^{12}_{6}\text{C}$. Taking this isotope as a typical carbon atom and using the data in Table 2.1, we can calculate the mass of one atom by summing the masses of six electrons, six protons, and six neutrons.

$$\begin{aligned}&6\ \text{electron}\ (9.10938\times10^{-28}\ \text{g/electron}) + 6\ \text{proton}\ (1.67262\times10^{-24}\ \text{g/proton})\\ &\quad + 6\ \text{neutron}\ (1.67493\times10^{-24}\ \text{g/neutron}) \approx 2\times10^{-23}\ \text{g}\end{aligned}$$

Therefore, the number of carbon atoms in a typical microsphere would be

$$3\times10^{-11}\ \text{g C} \times \frac{1\ \text{C atom}}{2\times10^{-23}\ \text{g C}} = 2\times10^{12}\ \text{C atoms}$$

This calculation gives us a feel for the incredibly small size of an atom. It should be noted that the mass of the carbon atom used in the calculation is only an approximation. Besides the assumption that all carbon atoms are carbon-12 (only about 99% of carbon atoms are carbon-12), we will see in Chapter 19 that the mass of an atom is actually somewhat less than the sum of the masses of the individual particles. In Chapter 3 we will introduce "atomic mass" which is the average mass of the naturally occurring isotopes of a particular element.

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