

FUNDAMENTALS

A.1 (a) chemical; (b) physical; (c) physical

A.3 The temperature, the humidity, and the evaporation of water are physical properties. The ripening of oranges is a chemical change.

A.5 (a) intensive; (b) intensive; (c) extensive; (d) extensive

A.7
$$d = \frac{m}{V}$$
$$= \left(\frac{112.32 \text{ g}}{29.27 \text{ mL} - 23.45 \text{ mL}} \right) \left(\frac{1 \text{ mL}}{1 \text{ cm}^3} \right)$$
$$= 19.3 \text{ g} \cdot \text{cm}^{-3}$$

A.9
$$d = \frac{m}{V}, \text{ rearranging gives } V = \frac{m}{d}$$
$$= \left(\frac{0.750 \text{ carat}}{3.51 \text{ g} \cdot \text{cm}^{-3}} \right) \left(\frac{200 \text{ mg}}{1 \text{ carat}} \right) \left(\frac{1 \text{ g}}{1000 \text{ mg}} \right)$$
$$= 0.0427 \text{ cm}^3$$

A.11
$$d = \frac{m}{V}$$
$$= \left(\frac{3.95 \times 10^{-22} \text{ g}}{\frac{4}{3} \pi (138 \text{ pm})^3} \right) \left(\frac{1 \text{ pm}}{1 \times 10^{-10} \text{ cm}} \right)^3$$
$$= 35.9 \text{ g} \cdot \text{cm}^{-3}$$

Because the density of metallic uranium is much less than the density of a uranium atom, the metallic form of uranium must contain considerable empty space.

$$\begin{aligned}
 \text{A.13 (a)} \quad d &= \frac{m}{V} \\
 &= \left(\frac{0.213 \text{ g}}{1.100 \text{ cm} \times 0.531 \text{ cm} \times 0.212 \text{ cm}} \right) \\
 &= \left(\frac{0.213 \text{ g}}{0.1238 \text{ cm}^3} \right) \\
 &= 1.72 \text{ g} \cdot \text{cm}^{-3}
 \end{aligned}$$

This determination is more precise because the volume is not limited to 2 significant figures as it is in part (b).

$$\begin{aligned}
 \text{(b)} \quad d &= \frac{m}{V} \\
 &= \left(\frac{41.003 \text{ g} - 39.753 \text{ g}}{20.37 \text{ mL} - 19.65 \text{ mL}} \right) \\
 &= \left(\frac{1.250 \text{ g}}{0.72 \text{ mL}} \right) \left(\frac{1 \text{ mL}}{1 \text{ cm}^3} \right) \\
 &= 1.7 \text{ g} \cdot \text{cm}^{-3}
 \end{aligned}$$

$$\begin{aligned}
 \text{A.15} \quad E_K &= \frac{1}{2}mv^2 \\
 &= \frac{1}{2}(4.2 \text{ kg})(14 \text{ km} \cdot \text{h}^{-1})^2 \left(\frac{1 \text{ h}}{3600 \text{ s}} \right)^2 \left(\frac{1000 \text{ m}}{1 \text{ km}} \right)^2 \\
 &= 32 \text{ kg} \cdot \text{m}^2 \cdot \text{s}^{-2} \\
 &= 32 \text{ J}
 \end{aligned}$$

$$\text{A.17} \quad m=2.8 \text{ metric tons}, \quad v_i = 100 \text{ km} \cdot \text{hr}^{-1}, \quad v_f = 50 \text{ km} \cdot \text{hr}^{-1}$$

$$E_K = \frac{1}{2}mv^2$$

$$\begin{aligned}
 E_{K(\text{init})} &= \frac{1}{2}(2.8 \text{ metric tons}) \left(\frac{10^3 \text{ kg}}{1 \text{ metric ton}} \right) \\
 &\quad \left[\left(\frac{100 \text{ km}}{1 \text{ hr}} \right) \left(\frac{1 \text{ hr}}{60 \text{ min.}} \right) \left(\frac{1 \text{ min.}}{60 \text{ sec.}} \right) \right]^2 \\
 &= 4.32 \text{ kg} \cdot \text{km}^2 \cdot \text{s}^{-2} = 4.32 \times 10^6 \text{ kg} \cdot \text{m}^2 \cdot \text{s}^{-2} = 4,320 \text{ kJ}
 \end{aligned}$$

$$E_{K(final)} = \frac{1}{2} (2.8 \text{ metric tons}) \left(\frac{10^3 \text{ kg}}{1 \text{ metric ton}} \right) \left[\left(\frac{50 \text{ km}}{1 \text{ hr}} \right) \left(\frac{1 \text{ hr}}{60 \text{ min.}} \right) \left(\frac{1 \text{ min.}}{60 \text{ sec.}} \right) \right]^2$$

$$= 0.27 \text{ kg} \cdot \text{km}^2 \cdot \text{s}^{-2} = 0.27 \times 10^6 \text{ kg} \cdot \text{m}^2 \cdot \text{s}^{-2} = 270 \text{ kJ}$$

$$E_{K(init)} - E_{K(final)} = (4,320 - 270) \text{ kJ} = 4,050 \text{ kJ} = 4.0 \times 10^3 \text{ kJ} \text{ (2 SF)}$$

This amount of energy could have been recovered, neglecting friction and other losses, or used to drive the vehicle up a hill.

$$E_p = mgh \quad g = 9.81 \text{ ms}^{-2}$$

Setting potential energy equal to 4,050 kJ = $4.05 \text{ kg m}^2 \text{ s}^{-2}$ and solving for height gives

$$h = \frac{E_p}{mg} = \left(\frac{4.05 \times 10^6 \text{ kg} \cdot \text{m}^2 \cdot \text{s}^{-2}}{(2800 \text{ kg})(9.81 \text{ ms}^{-2})} \right) = 147 \text{ m} = 150 \text{ m} \text{ (2 SF)}$$

A.19 $E_p = mgh$

$$= (40.0 \text{ g})(9.81 \text{ m} \cdot \text{s}^{-2})(0.50 \text{ m}) \left(\frac{1 \text{ kg}}{1000 \text{ g}} \right)$$

$$= 0.20 \text{ kg} \cdot \text{m}^2 \cdot \text{s}^{-2} \text{ for one raise of a fork.}$$

$$\text{For 30 raises, } (30)(0.20 \text{ kg} \cdot \text{m}^2 \cdot \text{s}^{-2}) = 6.0 \text{ J}$$

A.21 (a) The energy is all potential energy before the ball is dropped. After the ball has fallen halfway, half of the energy has been converted to kinetic energy.

$$E_K = \frac{1}{2} mgh$$

$$= \frac{1}{2} (0.95 \text{ kg})(9.81 \text{ m} \cdot \text{s}^{-2})(13.9 \text{ m})$$

$$= 65 \text{ kg} \cdot \text{m}^2 \cdot \text{s}^{-2}$$

$$= 65 \text{ J}$$

(b) When the ball hits the floor, all of the energy has been converted to kinetic energy.

$$\begin{aligned}
E_K &= mgh \\
&= (0.95 \text{ kg})(9.81 \text{ m} \cdot \text{s}^{-2})(13.9 \text{ m}) \\
&= 1.3 \times 10^2 \text{ kg} \cdot \text{m}^2 \cdot \text{s}^{-2} \\
&= 1.3 \times 10^2 \text{ J}
\end{aligned}$$

A.23 We need to use the expansion given in Exercise A.22 to help solve this problem. We also need to recognize that $E_p = egh$ for the small difference in distance, h , can be represented by subtracting E_p at distance r between the proton and electron from E_p at distance $r+h$.

$$E_p = \frac{e^2 h}{4\pi\epsilon_0 r^2} = egh$$

$$\text{So } g = \frac{e^2 h}{4\pi\epsilon_0 r^2} \left(\frac{1}{eh} \right) = \frac{e}{4\pi\epsilon_0 r^2} \text{ when } E_p = egh.$$

A.25 The relationship between distance of separation and potential energy for charged particles is given in section A.2, equation 4.

$$\begin{aligned}
E_p &= V(r) = \frac{q_1 q_2}{4\pi\epsilon_0 r} = \frac{(-e)(+e)}{4\pi\epsilon_0 r} \\
&= \frac{-(1.602 \times 10^{-19} \text{ C})^2}{4\pi(8.85419 \times 10^{-12} \text{ C}^2 \cdot \text{J}^{-1} \cdot \text{m}^{-1})(53 \times 10^{-12} \text{ m})} \\
&= -4.352 \times 10^{-18} \text{ J} \left(\frac{1 \text{ eV}}{1.602 \times 10^{-19} \text{ J}} \right) = -27.17 \text{ eV} = -27 \text{ eV (2 SF)}
\end{aligned}$$

Considering the proton and electron beginning at rest and at infinite separation sets the initial total energy to 0. Since the electron is not at rest in a hydrogen atom, its total energy is represented by equation 5:

$$E = E_K + E_p$$

We have only calculated the potential energy. The discrepancy between the calculated value of the potential energy, -27.7 eV, and the measured amount released, -13.6 eV, is the kinetic energy of the electron, 13.6 eV.

A.27 SI unit of pressure = Pa = $\text{N} \cdot \text{m}^{-2} = \left(\frac{\text{kg} \cdot \text{m}}{\text{s}^2} \right) \left(\frac{1}{\text{m}^2} \right) = \frac{\text{kg}}{\text{m} \cdot \text{s}^2}$

SI unit of volume = m^3 or ($\text{L} = \text{dm}^3 = 1 \times 10^{-3} \text{ m}^3$)

SI unit of energy = J = $\frac{\text{kg} \cdot \text{m}^2}{\text{s}^2}$

The product of pressure and volume is

$$1 \text{ Pa} \times 1 \text{ m}^3 = \left(\frac{\text{kg}}{\text{m} \cdot \text{s}^2} \right) (\text{m}^3) = \frac{\text{kg} \cdot \text{m}^2}{\text{s}^2} = 1 \text{ J}$$

(Note: If one uses L instead of m^3 for volume, the same overall units result.)

A.29 $105.50 \text{ g} - 43.50 \text{ g} = 62.00 \text{ g} = m_{\text{H}_2\text{O}}$

$$d = \frac{m}{V} \quad V = \frac{m}{d} = \frac{m_{\text{H}_2\text{O}}}{d_{\text{H}_2\text{O}}} = \frac{62.00 \text{ g}}{0.9999 \text{ g} \cdot \text{cm}^{-3}} = 62.00 \text{ cm}^3$$

$$d_{\text{liquid}} = \left(\frac{96.75 \text{ g} - 43.50 \text{ g}}{62.00 \text{ cm}^3} \right) = 0.8589 \text{ g} \cdot \text{cm}^{-3}$$

B.1 number of beryllium atoms = $\frac{\text{mass of sample}}{\text{mass of one atom}}$

$$= \left(\frac{0.210 \text{ g}}{1.50 \times 10^{-26} \text{ kg} \cdot \text{atom}^{-1}} \right) \left(\frac{1 \text{ kg}}{1000 \text{ g}} \right)$$

$$= 1.40 \times 10^{22} \text{ atoms}$$

B.3 (a) Radiation may pass through a metal foil. (b) All light (electromagnetic radiation) travels at the same speed; the slower speed supports the particle model. (c) This observation supports the radiation model. (d) This observation supports the particle model; electromagnetic radiation has no mass and no charge.

B.5 (a) 4.80×10^{-10} esu; (b) 14 electrons

- B.7** (a) 5p, 6n, 5e; (b) 5p, 5n, 5e; (c) 15p, 16n, 15e; (d) 92p, 146n, 92e
- B.9** (a) ^{194}Ir ; (b) ^{22}Ne ; (c) ^{51}V
- B.11** (a) they all have the same mass; (b) they have differing numbers of protons, neutrons, and electrons
- B.13** bromine-79, bromine-81
- B.15** (a) 0.5359; (b) 0.4639; (c) 0.0002526; (d) 463.9 kg
- B.17** (a) Rubidium, Group 1 metal; (b) Radium, Group 2 metal; (c) Ruthenium, Group 8 metal; (d) Radon, Group 18 nonmetal.
- B.19** (a) Os, metal; (b) Tl, metal; (c) At, nonmetal
- B.21** Fluorine, F, $Z=9$, gas; Chlorine, Cl, $Z=17$, gas; Bromine, Br, $Z=35$, liquid; Iodine, I, $Z=53$, solid
- B.23** (a) *d* block; (b) *s* block; (c) *p* block; (d) *d* block; (e) *p* block; (f) *d* block
- C.1** (a) An ionic compound is made up of ions. Sodium chloride is an example of an ionic compound. (b) A molecular compound is made up of molecules. Sucrose (sugar) is an example of a molecular compound. Ionic compounds have higher melting points than solids made of molecules. Ionic compounds dissolve only in polar solvents if at all. Molecules are more likely to dissolve in nonpolar solvents. See Chapter 5.

- C.3** The empirical formula can be represented as $(C_6H_{10}O_1)_n$. Since each molecule has 2 oxygen atoms, $n=2$ and the molecular formula is $C_{12}H_{20}O_2$.
- C.5** (a) Cesium is a metal in Group 1; it will form Cs^+ ions. (b) Iodine is a nonmetal in Group 17/VII and will form I^- ions. (c) Selenium is a Group 16/VII nonmetal and will form Se^{2-} ions. (d) Calcium is a Group 2 metal and will form Ca^{2+} ions.
- C.7** (a) $^4He^{2+}$ has 2 protons, 2 neutrons, and no electrons. (b) $^{15}N^{3-}$ has 7 protons, 8 neutrons, and 10 electrons. (c) $^{127}I^-$ has 53 protons, 74 neutrons, and 54 electrons. (d) $^{80}Se^{2-}$ has 34 protons, 46 neutrons, and 36 electrons.
- C.9** (a) $^{19}F^-$; (b) $^{24}Mg^{2+}$; (c) $^{128}Te^{2-}$; (d) $^{86}Rb^+$
- C.11** (a) Aluminum forms Al^{3+} ions; tellurium forms Te^{2-} ions. Two aluminum atoms produce a charge of $2 \times +3 = +6$. Three tellurium atoms produce a charge of $3 \times -2 = -6$. The formula for aluminum telluride is Al_2Te_3 . (b) Magnesium forms Mg^{2+} ions and oxygen forms O^{2-} ions. A magnesium ion produces a charge of +2, which is required to balance the charge on one O^{2-} ion. The formula for magnesium oxide is MgO . (c) Sodium forms +1 ions; sulfur forms -2 ions. The formula for sodium sulfide is Na_2S . (d) Rubidium forms +1 ions and iodine forms -1 ions. One iodide ions are required to balance the charge of one rubidium ion, so the formula is RbI .

C.13 (a) HCl, molecular compound (in the gas phase); (b) S₈, element (molecular substance); (c) CoS, ionic compound; (d) Ar, element; (e) CS₂, molecular compound; (f) SrBr₂, ionic compound

C.15 (a) Group 13; (b) aluminum, Al

C.17 The formula is Al₂O₃.

$$\begin{aligned}d &= \frac{m}{V} \\&= \left(\frac{102 \text{ g}}{2.5 \text{ cm} \times 3.0 \text{ cm} \times 4.0 \text{ cm}} \right) \\&= 3.4 \text{ g} \cdot \text{cm}^{-3}\end{aligned}$$

D.1 (a) Al₂O₃. Aluminum forms 3+ ions and oxygen forms 2− ions. (b) Strontium forms +2 ions and the phosphate ion is PO₄^{3−}, so the formula is Sr₃(PO₄)₂. (c) Aluminum forms +3 ions and the carbonate ion is CO₃^{2−}, giving a formula of Al₂(CO₃)₃. (d) Lithium forms Li⁺ ions and the nitride ion is N^{3−}. The formula of lithium nitride is Li₃N.

D.3 (a) calcium phosphate; (b) tin(II) fluoride, stannous fluoride; (c) vanadium(V) oxide; (d) copper(I) oxide, cuprous oxide

D.5 (a) TiO₂; (b) SiCl₄; (c) CS₂; (d) SF₄; (e) Li₂S; (f) SbF₅; (g) N₂O₅; (h) IF₇

D.7 (a) sulfur hexafluoride; (b) dinitrogen pentoxide; (c) nitrogen triiodide; (d) xenon tetrafluoride; (e) arsenic tribromide; (f) chlorine dioxide

D.9 (a) hydrochloric acid; (b) sulfuric acid; (c) nitric acid; (d) acetic acid; (e) sulfurous acid; (f) phosphoric acid

D.11 (a) Na_2O ; (b) K_2SO_4 ; (c) AgF ; (d) $\text{Zn}(\text{NO}_3)_2$; (e) Al_2S_3

D.13 (a) sodium sulfite; (b) iron(III) oxide or ferric oxide; (c) iron(II) oxide or ferrous oxide; (d) magnesium hydroxide; (e) nickel(II) sulfate hexahydrate; (f) phosphorus(V) chloride; (g) chromium(III) dihydrogen phosphate; (h) diarsenic trioxide; (i) ruthenium(II) chloride

D.15 (a) heptane; (b) propane; (c) pentane; (d) butane

D.17 (a) cobalt (III) oxide monohydrate; $\text{Co}_2\text{O}_3 \cdot \text{H}_2\text{O}$; (b) cobalt (II) hydroxide; $\text{Co}(\text{OH})_2$

D.19 $\text{E}=\text{C}$; methane; sodium carbide

D.21 (a) lithium aluminum hydride, ionic (with a molecular anion); (b) sodium hydride, ionic

D.23 (a) telluric acid; (b) sodium arsenate; (c) calcium selenite; (d) barium antimonate; (e) arsenic acid; (f) cobalt(III) tellurate

D.25 (a) butanoic acid; (b) 2,3-dimethylbutane; (c) 1,1-dichloropentane

E.1 (a) moles of people =
$$\frac{6.0 \times 10^9 \text{ people}}{6.022 \times 10^{23} \text{ people} \cdot \text{mol}^{-1}}$$
$$= 1.0 \times 10^{-14} \text{ mol}$$

$$\begin{aligned}
 \text{(b) time} &= \frac{1 \text{ mol peas}}{1.0 \times 10^{-14} \text{ mol} \cdot \text{s}^{-1}} = 1.0 \times 10^{14} \text{ s} \\
 (1.0 \times 10^{14} \text{ s}) &\left(\frac{1 \text{ h}}{3600 \text{ s}} \right) \left(\frac{1 \text{ day}}{24 \text{ h}} \right) \left(\frac{1 \text{ yr}}{365 \text{ days}} \right) = 3.2 \times 10^6 \text{ years}
 \end{aligned}$$

E.3 (a) mass of average Li atom

$$\begin{aligned}
 &= \left(\frac{7.42}{100} \right) (9.988 \times 10^{-24} \text{ g}) + \left(\frac{92.58}{100} \right) (1.165 \times 10^{-23} \text{ g}) \\
 &= 1.153 \times 10^{-23} \text{ g} \cdot \text{atom}^{-1}
 \end{aligned}$$

$$\begin{aligned}
 \text{molar mass} &= (1.153 \times 10^{-23} \text{ g} \cdot \text{atom}^{-1}) (6.022 \times 10^{23} \text{ atoms} \cdot \text{mol}^{-1}) \\
 &= 6.94 \text{ g} \cdot \text{mol}^{-1}
 \end{aligned}$$

(b) mass of average Li atom

$$\begin{aligned}
 &= \left(\frac{5.67}{100} \right) (9.988 \times 10^{-24} \text{ g}) + \left(\frac{100 - 5.67}{100} \right) (1.165 \times 10^{-23} \text{ g}) \\
 &= 1.1556 \times 10^{-23} \text{ g} \cdot \text{atom}^{-1}
 \end{aligned}$$

$$\begin{aligned}
 \text{molar mass} &= (1.1556 \times 10^{-23} \text{ g} \cdot \text{atom}^{-1}) (6.022 \times 10^{23} \text{ atoms} \cdot \text{mol}^{-1}) \\
 &= 6.96 \text{ g} \cdot \text{mol}^{-1}
 \end{aligned}$$

E.5 (a) $\text{MgSO}_4 \cdot 7 \text{ H}_2\text{O}$ formula mass = $246.48 \text{ g} \cdot \text{mol}^{-1}$

$$\text{atoms of O} = \left(\frac{5.15 \text{ g}}{246.48 \text{ g} \cdot \text{mol}^{-1}} \right) \left(\frac{11 \text{ mol O atoms}}{\text{mol MgSO}_4 \cdot 7 \text{ H}_2\text{O}} \right)$$

$$(6.022 \times 10^{23} \text{ atoms} \cdot \text{mol}^{-1}) = 1.38 \times 10^{23}$$

$$\begin{aligned}
 \text{(b) formula units} &= \left(\frac{5.15 \text{ g}}{246.48 \text{ g} \cdot \text{mol}^{-1}} \right) (6.022 \times 10^{23} \text{ atoms} \cdot \text{mol}^{-1}) \\
 &= 1.26 \times 10^{22}
 \end{aligned}$$

$$\text{(c) moles of H}_2\text{O} = 7 \left(\frac{5.15 \text{ g}}{246.48 \text{ g} \cdot \text{mol}^{-1}} \right) = 0.146 \text{ mol}$$

E.7 The percentage $^{10}\text{B} = 100 - \text{percentage } ^{11}\text{B}$

$$\begin{aligned}\text{molar mass} &= \left(\frac{\% \text{ } ^{10}\text{B}}{100\%} \right) (\text{mass } ^{10}\text{B}) + \left(\frac{\% \text{ } ^{11}\text{B}}{100\%} \right) (\text{mass } ^{11}\text{B}) \\ &= \left(\frac{100\% - \% \text{ } ^{11}\text{B}}{100\%} \right) (\text{mass } ^{10}\text{B}) + \left(\frac{\% \text{ } ^{11}\text{B}}{100\%} \right) (\text{mass } ^{11}\text{B})\end{aligned}$$

Rearranging gives

$$\begin{aligned}\% \text{ } ^{11}\text{B} &= \frac{100 \cdot \text{molar mass} - 100 \cdot \text{mass } ^{10}\text{B}}{\text{mass } ^{11}\text{B} - \text{mass } ^{10}\text{B}} \\ &= \frac{100(10.81 \text{ g} \cdot \text{mol}^{-1}) - 100(10.013 \text{ g} \cdot \text{mol}^{-1})}{11.093 \text{ g} \cdot \text{mol}^{-1} - 10.013 \text{ g} \cdot \text{mol}^{-1}} \\ &= 73.8 \% \\ \% \text{ } ^{10}\text{B} &= 26.2 \%\end{aligned}$$

E.9 (a) $\frac{75 \text{ g}}{114.82 \text{ g} \cdot \text{mol}^{-1}} = 0.65 \text{ mol In}$

$\frac{80 \text{ g}}{127.60 \text{ g} \cdot \text{mol}^{-1}} = 0.63 \text{ mol Te}$

80. g of tellurium contains more moles of atoms than does 75 g of indium.

(b) $\frac{15.0 \text{ g}}{30.97 \text{ g} \cdot \text{mol}^{-1}} = 0.484 \text{ mol P}$

$\frac{15.0 \text{ g}}{32.07 \text{ g} \cdot \text{mol}^{-1}} = 0.468 \text{ mol S}$

15.0 g of P has slightly more atoms than 15.0 g of S.

(c) Because the two samples have the same number of atoms, they will have the same number of moles, which is given by

$$\frac{2.49 \times 10^{22} \text{ atoms}}{6.022 \times 10^{23} \text{ atoms} \cdot \text{mol}^{-1}} = 0.0413 \text{ mol}$$

E.11 (a) $m_{\text{Rh}} = \left(\frac{57 \text{ g N}}{14.01 \text{ g} \cdot \text{mol}^{-1} \text{ N}} \right) (102.91 \text{ g} \cdot \text{mol}^{-1} \text{ Rh})$

$= 4.1 \times 10^2 \text{ g Rh}$

$$\begin{aligned} \text{(b) } m_{\text{Rh}} &= \left(\frac{57 \text{ g Zr}}{91.22 \text{ g} \cdot \text{mol}^{-1} \text{ Zr}} \right) (102.91 \text{ g} \cdot \text{mol}^{-1} \text{ Rh}) \\ &= 63 \text{ g Rh} \end{aligned}$$

E.13 (a) molar mass of $\text{Al}_2\text{O}_3 = 101.96 \text{ g} \cdot \text{mol}^{-1}$

$$n_{\text{Al}_2\text{O}_3} = \frac{10.0 \text{ g}}{101.96 \text{ g} \cdot \text{mol}^{-1}} = 0.0981 \text{ mol}$$

$$N_{\text{Al}_2\text{O}_3} = (0.0981 \text{ mol})(6.022 \times 10^{23} \text{ atoms} \cdot \text{mol}^{-1}) = 5.91 \times 10^{22} \text{ molecules}$$

(b) molar mass of $\text{HF} = 20.01 \text{ g} \cdot \text{mol}^{-1}$

$$n_{\text{HF}} = \frac{25.92 \times 10^{-3} \text{ g}}{20.01 \text{ g} \cdot \text{mol}^{-1}} = 1.30 \times 10^{-3} \text{ mol}$$

$$\begin{aligned} N_{\text{HF}} &= (1.30 \times 10^{-3} \text{ mol})(6.022 \times 10^{23} \text{ atoms} \cdot \text{mol}^{-1}) \\ &= 7.83 \times 10^{20} \text{ molecules} \end{aligned}$$

(c) molar mass of hydrogen peroxide = $34.02 \text{ g} \cdot \text{mol}^{-1}$

$$n_{\text{H}_2\text{O}_2} = \frac{1.55 \times 10^{-3} \text{ g}}{34.02 \text{ g} \cdot \text{mol}^{-1}} = 4.56 \times 10^{-5} \text{ mol}$$

$$\begin{aligned} N_{\text{H}_2\text{O}_2} &= (4.56 \times 10^{-5} \text{ mol})(6.022 \times 10^{23} \text{ atoms} \cdot \text{mol}^{-1}) \\ &= 2.75 \times 10^{19} \text{ molecules} \end{aligned}$$

(d) molar mass of glucose = $180.15 \text{ g} \cdot \text{mol}^{-1}$

$$n_{\text{glucose}} = \frac{1250 \text{ g}}{180.15 \text{ g} \cdot \text{mol}^{-1}} = 6.94 \text{ mol}$$

$$N_{\text{glucose}} = (6.94 \text{ mol})(6.022 \times 10^{23} \text{ atoms} \cdot \text{mol}^{-1}) = 4.18 \times 10^{24} \text{ molecules}$$

(e) molar mass of N atoms = $14.01 \text{ g} \cdot \text{mol}^{-1}$

$$n_{\text{N}} = \frac{4.37 \text{ g}}{14.01 \text{ g} \cdot \text{mol}^{-1}} = 0.312 \text{ mol}$$

$$N_{\text{N}} = (0.312 \text{ mol})(6.022 \times 10^{23} \text{ atoms} \cdot \text{mol}^{-1}) = 1.88 \times 10^{23} \text{ atoms}$$

molar mass of N_2 molecules = $28.02 \text{ g} \cdot \text{mol}^{-1}$

$$n_{\text{N}_2} = \frac{4.37 \text{ g}}{28.02 \text{ g} \cdot \text{mol}^{-1}} = 0.156 \text{ mol}$$

$$N_{\text{N}_2} = (0.156 \text{ mol})(6.022 \times 10^{23} \text{ atoms} \cdot \text{mol}^{-1}) = 9.39 \times 10^{22} \text{ molecules}$$

E.15 (a) molar mass of $\text{AgCl} = 143.32 \text{ g} \cdot \text{mol}^{-1}$

$$n_{\text{AgCl}} = \frac{2.00 \text{ g}}{143.32 \text{ g} \cdot \text{mol}^{-1}} = 0.0140 \text{ mol}$$

The number of moles of Ag^+ ions equals the number of moles of AgCl .

(b) molar mass of $\text{UO}_3 = 286.03 \text{ g} \cdot \text{mol}^{-1}$

$$n_{\text{UO}_3} = \frac{600 \text{ g}}{286.03 \text{ g} \cdot \text{mol}^{-1}} = 2.10 \text{ mol}$$

(c) molar mass of $\text{FeCl}_3 = 162.20$

$$n_{\text{FeCl}_3} = \left(\frac{4.19 \text{ mg}}{162.20 \text{ g} \cdot \text{mol}^{-1}} \right) \left(\frac{1 \text{ g}}{1000 \text{ mg}} \right) = 2.58 \times 10^{-5} \text{ mol}$$

The number of moles of Cl^- ions equals 3 times the number of moles of FeCl_3 .

$$n_{\text{Cl}^-} = 7.75 \times 10^{-5} \text{ mol}$$

(d) molar mass of $\text{AuCl}_3 \cdot 2 \text{ H}_2\text{O} = 339.35 \text{ g} \cdot \text{mol}^{-1}$

$$\begin{aligned} \left(\frac{2 \text{ mol H}_2\text{O}}{1 \text{ mol AuCl}_3 \cdot 2 \text{ H}_2\text{O}} \right) (n_{\text{AuCl}_3 \cdot 2 \text{ H}_2\text{O}}) &= 2 \left(\frac{1.00 \text{ g}}{339.35 \text{ g} \cdot \text{mol}^{-1}} \right) \\ &= 5.89 \times 10^{-3} \text{ mol H}_2\text{O} \end{aligned}$$

E.17 (a) number of formula units = $(0.750 \text{ mol})(6.022 \times 10^{23} \text{ formula units} \cdot \text{mol}^{-1})$

$$= 4.52 \times 10^{23} \text{ formula units}$$

(b) molar mass of $\text{Ag}_2\text{SO}_4 = 311.80 \text{ g} \cdot \text{mol}^{-1}$

$$\begin{aligned} \left(\frac{2.39 \times 10^{20} \text{ formula units}}{6.022 \times 10^{23} \text{ formula units} \cdot \text{mol}^{-1}} \right) (311.80 \text{ g} \cdot \text{mol}^{-1}) \left(\frac{1000 \text{ mg}}{1 \text{ g}} \right) \\ = 124 \text{ mg} \end{aligned}$$

(c) molar mass of $\text{NaHCO}_2 = 68.01 \text{ g} \cdot \text{mol}^{-1}$

$$\left(\frac{3.429 \text{ g}}{68.01 \text{ g} \cdot \text{mol}^{-1}} \right) (6.022 \times 10^{23} \text{ formula units} \cdot \text{mol}^{-1})$$

$$= 3.036 \times 10^{22} \text{ formula units}$$

E.19 (a) molar mass of $\text{H}_2\text{O} = 18.02 \text{ g} \cdot \text{mol}^{-1}$

$$\left(\frac{18.02 \text{ g} \cdot \text{mol}^{-1}}{6.022 \times 10^{23} \text{ molecules} \cdot \text{mol}^{-1}} \right) = 2.992 \times 10^{-23} \text{ g} \cdot \text{molecule}^{-1}$$

$$(b) \ N_{\text{H}_2\text{O}} = \left(\frac{1000 \text{ g}}{18.02 \text{ g} \cdot \text{mol}^{-1}} \right) (6.022 \times 10^{23} \text{ molecules} \cdot \text{mol}^{-1})$$

$$= 3.34 \times 10^{25} \text{ molecules}$$

E.21 (a) molar mass of $\text{CuBr}_2 \cdot 4 \text{H}_2\text{O} = 295.42 \text{ g} \cdot \text{mol}^{-1}$

$$\left(\frac{7.35 \text{ g}}{295.42 \text{ g} \cdot \text{mol}^{-1}} \right) = 2.49 \times 10^{-2} \text{ mol}$$

(b) Because there are 2 mol Br^- per mole of compound, the number of moles will be twice the amount in part (a), $4.98 \times 10^{-2} \text{ mol}$.

(c)

$$\left(\frac{4 \text{ mol H}_2\text{O}}{1 \text{ mol CuBr}_2 \cdot 4 \text{H}_2\text{O}} \right) (2.49 \times 10^{-2} \text{ mol}) (6.022 \times 10^{23} \text{ molecules} \cdot \text{mol}^{-1})$$

$$= 6.00 \times 10^{22} \text{ molecules H}_2\text{O}$$

$$(d) \text{ fraction of mass due to O} = \frac{4(16.00 \text{ g} \cdot \text{mol}^{-1})}{295.42 \text{ g} \cdot \text{mol}^{-1}} = 0.2166$$

E.23 (a)

$$\left(\frac{10 \text{ mol H}_2\text{O}}{1 \text{ mol hydrated cpd}} \right) \left(\frac{18.01 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} \right) \left(\frac{1 \text{ mol hydrated cpd}}{264.166 \text{ g hydrated cpd}} \right) \times 100$$

$$= 68.2\% \text{ H}_2\text{O}$$

Therefore 6.82 kg out of 10 kg was water.

$$\frac{? \$}{\text{L H}_2\text{O}} = \left(\frac{\$72.00}{10 \text{ kg hydrated cpd}} \right) \left(\frac{10 \text{ kg hydrated cpd}}{6.82 \text{ kg H}_2\text{O}} \right) \left(\frac{1 \text{ kg H}_2\text{O}}{1 \text{ L H}_2\text{O}} \right)$$

$$= \$10.60 \text{ per L H}_2\text{O}$$

(b) Since the anhydrous compound costs \$80.00 for 10 kg, or \$8.00/kg, and only 10-6.82=3.18 kg is NaHCO₃, a fair price would have been

$$\frac{\$8.00}{\text{kg}} \times 3.18 \text{ kg} = \$25.44.$$

E.25 (a) moles of Cu = $\frac{43.4 \text{ g}}{63.54 \text{ g} \cdot \text{mol}^{-1}} = 0.683 \text{ mol}$

$$\text{S atoms required} = (0.683 \text{ mol})(6.022 \times 10^{23} \text{ atoms} \cdot \text{mol}^{-1}) = 4.11 \times 10^{23}$$

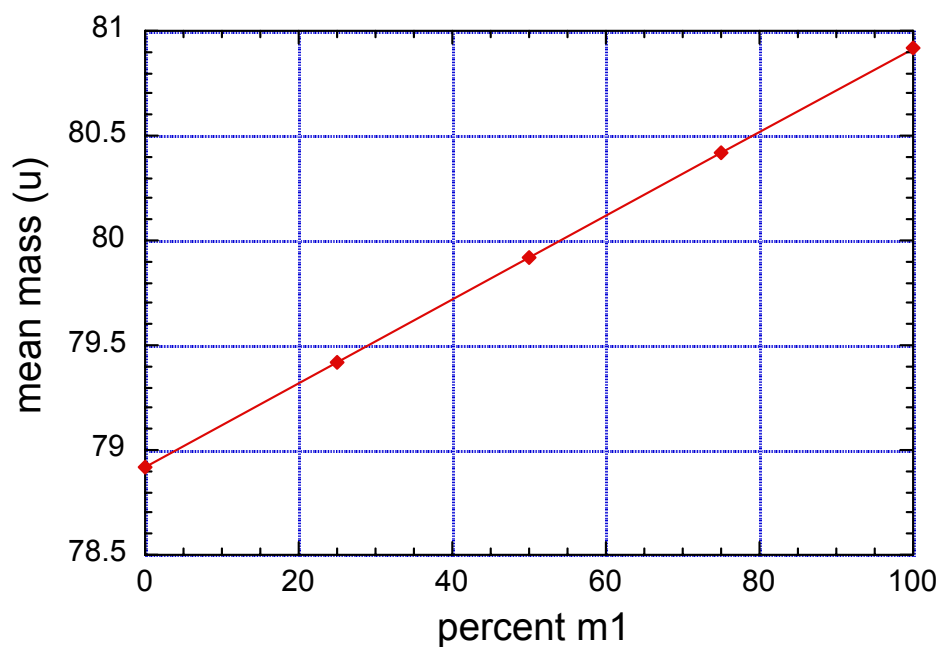
(b) S₈ molecules required = $\frac{4.11 \times 10^{23}}{8} = 5.14 \times 10^{22}$

(c) mass of sulfur needed = $(0.683 \text{ mol})(32.06 \text{ g} \cdot \text{mol}^{-1}) = 21.9 \text{ g}$

E.27 Solve by factor label (dimensional analysis).

$$\begin{aligned} ? \text{ mole H atoms} &= 28.0 \text{ cm}^3 \text{ NaBH}_4 \left(\frac{1.074 \text{ g}}{1 \text{ cm}^3} \right) \left(\frac{2.50 \times 10^{23} \text{ H atoms}}{3.93 \text{ g}} \right) \\ &\quad \times \left(\frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ H atoms}} \right) \\ &= 3.18 \text{ moles H atoms} \end{aligned}$$

E.29 For example, Br has m₁ = 80.9163 u, m₂ = 78.9183 u



F.1 $\text{C}_7\text{H}_{15}\text{NO}_3$

$$7 \times 12.01 \text{ g} = 84.07 \text{ g} \quad \times \left(\frac{100}{161.20 \text{ g}} \right) = 52.15\%$$

$$15 \times 1.0079 \text{ g} = 15.1185 \text{ g} \quad \times \left(\frac{100}{161.20 \text{ g}} \right) = 9.3787\%$$

$$1 \times 14.01 \text{ g} = 14.01 \text{ g} \quad \times \left(\frac{100}{161.20 \text{ g}} \right) = 8.691\%$$

$$3 \times 16.00 \text{ g} = \underline{48.00 \text{ g}} \quad \times \left(\frac{100}{161.20 \text{ g}} \right) = \underline{29.78\%}$$

161.20 g 100.00%

F.3 (a) M_2O 88.8% M For 100 g of compound, 88.8 g is M, 11.2 g is O.

$$\frac{? \text{ g M}_2\text{O}}{\text{mole M}_2\text{O}} = \left(\frac{100 \text{ g M}_2\text{O}}{11.2 \text{ g O}} \right) \left(\frac{16.00 \text{ g O}}{1 \text{ mol O}} \right) \left(\frac{1 \text{ mol O}}{1 \text{ mol M}_2\text{O}} \right)$$

$$= 143 \text{ g} \cdot \text{mol}^{-1} \text{ M}_2\text{O}$$

Therefore, $143 - 16 = 127 \text{ g/mol}$ are due to M. Since there are 2 moles of M per mole of M_2O ,

the molar mass of M=63.4 g/mol. That molar mass matches Cu. (b)
copper (I) oxide

F.5 (a) For 100 g of compound,

$$\text{moles of Na} = \frac{32.79 \text{ g}}{22.99 \text{ g} \cdot \text{mol}^{-1}} = 1.426 \text{ mol}$$

$$\text{moles of Al} = \frac{13.02 \text{ g}}{26.98 \text{ g} \cdot \text{mol}^{-1}} = 0.4826 \text{ mol}$$

$$\text{moles of F} = \frac{54.19 \text{ g}}{19.00 \text{ g} \cdot \text{mol}^{-1}} = 2.852 \text{ mol}$$

Dividing each number by 0.4826 gives a ratio of 1 Al : 2.95 Na : 5.91 F.

The formula is Na_3AlF_6 .

(b) For 100 g of compound,

$$\text{moles of K} = \frac{31.91 \text{ g}}{39.10 \text{ g} \cdot \text{mol}^{-1}} = 0.8161 \text{ mol}$$

$$\text{moles of Cl} = \frac{28.93 \text{ g}}{35.45 \text{ g} \cdot \text{mol}^{-1}} = 0.8161 \text{ mol}$$

mass of O is obtained by difference:

$$\text{moles of O} = \frac{100 \text{ g} - 31.91 \text{ g} - 28.93 \text{ g}}{16.00 \text{ g} \cdot \text{mol}^{-1}} = 2.448 \text{ mol}$$

Dividing each number by 0.8161 gives a ratio of 1.00 K : 1 Cl : 3.00 O.

The formula is KClO_3 .

(c) For 100 g of compound,

$$\text{moles of N} = \frac{12.2 \text{ g}}{14.01 \text{ g} \cdot \text{mol}^{-1}} = 0.871 \text{ mol}$$

$$\text{moles of H} = \frac{5.26 \text{ g}}{1.0079 \text{ g} \cdot \text{mol}^{-1}} = 5.22 \text{ mol}$$

$$\text{moles of P} = \frac{26.9 \text{ g}}{30.97 \text{ g} \cdot \text{mol}^{-1}} = 0.869 \text{ mol}$$

$$\text{moles of O} = \frac{55.6 \text{ g}}{16.00 \text{ g} \cdot \text{mol}^{-1}} = 3.475 \text{ mol}$$

Dividing each number by 0.869 gives a ratio of 1.00 N : 6.01 H : 1.00 P : 4.00 O. The formula is NH_6PO_4 or $[\text{NH}_4][\text{H}_2\text{PO}_4]$, ammonium dihydrogen phosphate.

F.7 moles of P = $\frac{4.14 \text{ g}}{30.97 \text{ g} \cdot \text{mol}^{-1}} = 0.134 \text{ mol}$

moles of Cl = $\frac{27.8 \text{ g} - 4.14 \text{ g}}{35.45 \text{ g} \cdot \text{mol}^{-1}} = 0.667 \text{ mol}$

Dividing each number by 0.134 mol gives a ratio of 4.98 Cl : 1 P. The formula is PCl_5 .

F.9 For 100 g of compound,

moles of C = $\frac{54.82 \text{ g}}{12.01 \text{ g} \cdot \text{mol}^{-1}} = 4.565 \text{ mol}$

moles of H = $\frac{5.62 \text{ g}}{1.0079 \text{ g} \cdot \text{mol}^{-1}} = 5.58 \text{ mol}$

moles of N = $\frac{7.10 \text{ g}}{14.01 \text{ g} \cdot \text{mol}^{-1}} = 0.507 \text{ mol}$

moles of O = $\frac{32.46 \text{ g}}{16.00 \text{ g} \cdot \text{mol}^{-1}} = 2.029 \text{ mol}$

Dividing each number by 0.507 mol gives a ratio of 9.00 C : 11.01 H : 1.00 N : 4.00 O. The formula is $\text{C}_9\text{H}_{11}\text{NO}_4$.

F.11 For 100 g of the osmium carbonyl compound,

moles of C = $\frac{15.89 \text{ g}}{12.01 \text{ g} \cdot \text{mol}^{-1}} = 1.323 \text{ mol}$

moles of O = $\frac{21.18 \text{ g}}{16.00 \text{ g} \cdot \text{mol}^{-1}} = 1.324 \text{ mol}$

moles of Os = $\frac{62.93 \text{ g}}{190.2 \text{ g} \cdot \text{mol}^{-1}} = 0.3309 \text{ mol}$

Dividing each number by 0.3309 mol gives a ratio of 4.00 C : 4.00 O : 1.00 Os. (a) The empirical formula is OsC_4O_4 . (b) The formula mass of OsC_4O_4 is $302.24 \text{ g} \cdot \text{mol}^{-1}$. The molar mass is $907 \text{ g} \cdot \text{mol}^{-1}$ which is 3 times the formula mass, so the molecular formula is $\text{Os}_3\text{C}_{12}\text{O}_{12}$.

F.13 For 100 g of caffeine,

$$\text{moles of C} = \frac{49.48 \text{ g}}{12.01 \text{ g} \cdot \text{mol}^{-1}} = 4.12 \text{ mol}$$

$$\text{moles of H} = \frac{5.19 \text{ g}}{1.0079 \text{ g} \cdot \text{mol}^{-1}} = 5.15 \text{ mol}$$

$$\text{moles of N} = \frac{28.85 \text{ g}}{14.01 \text{ g} \cdot \text{mol}^{-1}} = 2.059 \text{ mol}$$

$$\text{moles of O} = \frac{16.48 \text{ g}}{16.00 \text{ g} \cdot \text{mol}^{-1}} = 1.03 \text{ mol}$$

Dividing each number by 1.03 mol gives a ratio of 4.00 C : 5.00 H : 2.00 N : 1.00 O. The formula is $\text{C}_4\text{H}_5\text{N}_2\text{O}$ with a molar formula mass of $97.10 \text{ g} \cdot \text{mol}^{-1}$. Because the molecular molar mass is twice this value, the actual formula will be $\text{C}_8\text{H}_{10}\text{N}_4\text{O}_2$.

F.15 Glucose ($\text{C}_6\text{H}_{12}\text{O}_6$) has a molar mass of $180.15 \text{ g} \cdot \text{mol}^{-1}$ and will have the following composition:

$$\% \text{C} = \frac{6(12.01 \text{ g} \cdot \text{mol}^{-1})}{180.15 \text{ g} \cdot \text{mol}^{-1}} = 40.00\%$$

$$\% \text{H} = \frac{12(1.0079 \text{ g} \cdot \text{mol}^{-1})}{180.15 \text{ g} \cdot \text{mol}^{-1}} = 6.71\%$$

$$\% \text{O} = \frac{6(16.00 \text{ g} \cdot \text{mol}^{-1})}{180.15 \text{ g} \cdot \text{mol}^{-1}} = 53.29\%$$

Sucrose ($\text{C}_{12}\text{H}_{22}\text{O}_{11}$) has a molar mass of $342.29 \text{ g} \cdot \text{mol}^{-1}$ and will have the following composition:

$$\%C = \frac{12(12.01 \text{ g} \cdot \text{mol}^{-1})}{342.29 \text{ g} \cdot \text{mol}^{-1}} = 42.10\%$$

$$\%H = \frac{22(1.0079 \text{ g} \cdot \text{mol}^{-1})}{342.29 \text{ g} \cdot \text{mol}^{-1}} = 6.48\%$$

$$\%O = \frac{11(16.00 \text{ g} \cdot \text{mol}^{-1})}{342.29 \text{ g} \cdot \text{mol}^{-1}} = 51.42\%$$

While the %H values for glucose and sucrose are too close to allow us to distinguish between them by this value alone, %C (40.00 versus 42.10%) and %O (53.29 versus 51.42%) values are sufficient that, when taken together, can give us a reasonable amount of confidence in distinguishing between them.

F.17 Calculate the mass percent carbon for each fuel from its formula.

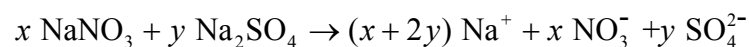
$$\text{ethene, } C_2H_4 \quad \frac{2(12.01)}{2(12.01)+4(1.0079)} \times 100 = 85.63\% \text{ C}$$

$$\text{propanol, } C_3H_7OH \quad \frac{3(12.01)}{3(12.01)+8(1.0079)+16.00} \times 100 = 59.96\% \text{ C}$$

$$\text{heptane, } C_7H_{16} \quad \frac{7(12.01)}{7(12.01)+16(1.0079)} \times 100 = 83.91\% \text{ C}$$

ethene (85.63%) > heptane (83.91%) > propanol (59.96%)

F.19 This problem requires that we relate unknowns to each other appropriately by writing a balanced chemical equation and using other information in the problem.



$$\frac{1.61 \text{ g Na}^+}{22.99 \text{ g} \cdot \text{mol}^{-1} \text{ Na}^+} = 0.07003 \text{ mol Na}^+ = x + 2y$$

$$5.37 \text{ g total} - 1.61 \text{ g Na}^+ = 3.76 \text{ g} = (62.01 \text{ g} \cdot \text{mol}^{-1})x + (96.07 \text{ g} \cdot \text{mol}^{-1})y$$

Rearrange and substitute:

$$3.76 = 62.01(0.07003 - 2y) + 96.07y$$

$$0.06065 = 0.07003 - 2y + 1.548y$$

$$0.009385 = 0.4516y$$

$$y = 0.02078 \text{ moles sulfate, } x = 0.02847 \text{ moles nitrate}$$

Therefore, the mass of sodium nitrate in the mixture was

$$0.02847 \text{ mol} \left(\frac{85.00 \text{ g NaNO}_3}{1 \text{ mol}} \right) = 2.42 \text{ g NaNO}_3$$

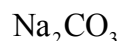
$$\frac{2.42 \text{ g NaNO}_3}{5.37 \text{ g total}} \times 100 = 45.1\% \text{ NaNO}_3$$

G.1 (a) solubility; (b) the abilities of the components to adsorb; (c) boiling points

G.3 (a) homogeneous, distillation; (b) heterogeneous, dissolving followed by filtration and distillation; (c) homogeneous, distillation

G.5 mass of $\text{AgNO}_3 = (0.179 \text{ mol} \cdot \text{L}^{-1})(0.5000 \text{ L})(169.88 \text{ g} \cdot \text{mol}^{-1}) = 15.2 \text{ g}$

G.7 (a) molarity of $\text{Na}_2\text{CO}_3 = \frac{2.111 \text{ g}}{(105.99 \text{ g} \cdot \text{mol}^{-1})(0.2500 \text{ L})} = 0.07967 \text{ M}$



$$V = \frac{(2.15 \times 10^{-3} \text{ mol Na}^+)(1 \text{ mol Na}_2\text{CO}_3)}{(0.07967 \text{ mol} \cdot \text{L}^{-1} \text{ Na}_2\text{CO}_3)(2 \text{ mol Na}^+)} = 1.35 \times 10^{-2} \text{ L or } 13.5 \text{ mL}$$

$$\begin{aligned} \text{(b)} \quad V &= \frac{(4.98 \times 10^{-3} \text{ mol CO}_3^{2-})(1 \text{ mol Na}_2\text{CO}_3)}{(0.07967 \text{ mol} \cdot \text{L}^{-1} \text{ Na}_2\text{CO}_3)(1 \text{ mol CO}_3^{2-})} \\ &= 6.25 \times 10^{-2} \text{ L or } 62.5 \text{ mL} \end{aligned}$$

$$\begin{aligned} \text{(c)} \quad V &= \frac{(50.0 \times 10^{-3} \text{ g Na}_2\text{CO}_3)}{(105.99 \text{ g} \cdot \text{mol}^{-1})(0.07967 \text{ mol} \cdot \text{L}^{-1} \text{ Na}_2\text{CO}_3)} \\ &= 5.92 \times 10^{-3} \text{ L or } 5.92 \text{ mL} \end{aligned}$$

G.9 (a) Weigh 1.6 g (0.010 mol, molar mass of $\text{KMnO}_4 = 158.04 \text{ g} \cdot \text{mol}^{-1}$) into a 1.0-L volumetric flask and add water to give a total volume of 1.0 L. Smaller (or larger) volumes could also be prepared by using a proportionally smaller (or larger) mass of KMnO_4 .

(b) Starting with $0.050 \text{ mol} \cdot \text{L}^{-1} \text{ KMnO}_4$, add four volumes of water to one volume of starting solution, because the concentration desired is one-fifth of the starting solution. This relation can be derived from the expression

$$V_i \times \text{molarity}_i = V_f \times \text{molarity}_f$$

where i represents the initial solution and f the final solution. But

$V_f = V_i + V_d$ where V_d represents the volume of solvent that must be added to dilute the initial solution. Rearranging the first equation gives

$$\frac{V_i}{V_f} = \frac{\text{molarity}_f}{\text{molarity}_i}$$

$$\frac{V_i}{V_i + V_d} = \frac{\text{molarity}_f}{\text{molarity}_i}$$

So if the ratio of final molarity to initial molarity is 1 : 5, we can write

$$\frac{V_i}{V_i + V_d} = \frac{1}{5}$$

$$5V_i = V_i + V_d$$

$$4V_i = V_d$$

For example, to prepare 50 mL of solution, you would add 40 mL of water to 10 mL of $0.050 \text{ mol} \cdot \text{L}^{-1} \text{ KMnO}_4$.

G.11 (a) $V(0.778 \text{ mol} \cdot \text{L}^{-1}) = (0.1500 \text{ L})(0.0234 \text{ mol} \cdot \text{L}^{-1})$

$$V = 4.51 \times 10^{-3} \text{ L or } 4.51 \text{ mL}$$

(b) The concentration desired is one-fifth of the starting NaOH solution, so the stockroom attendant will need to add four volumes of water to one volume of the $2.5 \text{ mol} \cdot \text{L}^{-1}$ solution. To prepare 60.0 mL of solution,

divide 60.0 by 5; so 12.0 mL of $2.5 \text{ mol} \cdot \text{L}^{-1}$ NaOH solution are added to 48.0 mL of water. (See the solution to G.9.)

G.13 (a) mass of $\text{CuSO}_4 = (0.20 \text{ mol} \cdot \text{L}^{-1})(0.250 \text{ L})(159.60 \text{ g} \cdot \text{mol}^{-1})$
 $= 8.0 \text{ g}$

(b) mass of $\text{CuSO}_4 \cdot 5 \text{ H}_2\text{O} = (0.20 \text{ mol} \cdot \text{L}^{-1})(0.250 \text{ L})(249.68 \text{ g} \cdot \text{mol}^{-1})$
 $= 12 \text{ g}$

G.15 (a) Chloride ions are supplied only by the $\text{NiCl}_2 \cdot 6 \text{ H}_2\text{O}$ complex:

$$[\text{Cl}^-] = \frac{(2 \text{ mol Cl}^-)(0.129 \text{ g NiCl}_2 \cdot 6 \text{ H}_2\text{O})}{(1 \text{ mol NiCl}_2 \cdot 6 \text{ H}_2\text{O})(237.70 \text{ g} \cdot \text{mol}^{-1} \text{ NiCl}_2 \cdot 6 \text{ H}_2\text{O})(0.250 \text{ L})}$$

$$= 0.00434 \text{ M}$$

(b) Ni^{2+} ions are present in both the $\text{NiSO}_4 \cdot 6 \text{ H}_2\text{O}$ and the $\text{NiCl}_2 \cdot 6 \text{ H}_2\text{O}$, so the final concentration will be the sum of the ions provided from the two sources:

Ni^{2+} from $\text{NiCl}_2 \cdot 6 \text{ H}_2\text{O}$

$$m_{\text{Ni}^{2+} \text{ from nickel chloride}} = \frac{(1 \text{ mol Ni}^{2+})(0.129 \text{ g NiCl}_2 \cdot 6 \text{ H}_2\text{O})}{(1 \text{ mol NiCl}_2 \cdot 6 \text{ H}_2\text{O})(237.70 \text{ g} \cdot \text{mol}^{-1} \text{ NiCl}_2 \cdot 6 \text{ H}_2\text{O})}$$

$$= 0.000543 \text{ mol}$$

Ni^{2+} from $\text{NiSO}_4 \cdot 6 \text{ H}_2\text{O}$

$$m_{\text{Ni}^{2+} \text{ from nickel chloride}} = \frac{(1 \text{ mol Ni}^{2+})(0.376 \text{ g NiSO}_4 \cdot 6 \text{ H}_2\text{O})}{(1 \text{ mol NiSO}_4 \cdot 6 \text{ H}_2\text{O})(262.86 \text{ g} \cdot \text{mol}^{-1} \text{ NiSO}_4 \cdot 6 \text{ H}_2\text{O})}$$

$$= 0.00143 \text{ mol}$$

$$\text{total moles of Ni}^{2+} = 0.000543 \text{ mol} + 0.00143 \text{ mol} = 0.00197 \text{ mol}$$

$$[\text{Ni}^{2+}] = \frac{0.00197 \text{ mol}}{0.2500 \text{ L}} = 0.00789 \text{ M}$$

G.17 (a) mass of $\text{K}_2\text{SO}_4 = (0.125 \text{ mol} \cdot \text{L}^{-1})(1.00 \text{ L})(174.26 \text{ g} \cdot \text{mol}^{-1}) = 21.8 \text{ g}$

$$(b) \text{ mass of NaF} = (0.015 \text{ mol} \cdot \text{L}^{-1})(0.375 \text{ L})(41.99 \text{ g} \cdot \text{mol}^{-1}) = 0.24 \text{ g}$$

$$(c) \text{ mass of C}_{12}\text{H}_{22}\text{O}_{11} = (0.35 \text{ mol} \cdot \text{L}^{-1})(0.500 \text{ L})(342.29 \text{ g} \cdot \text{mol}^{-1}) \\ = 60. \text{ g}$$

G.19 We can show that fewer than 1 molecule of X would be left after only 70 doublings.

$$10 \text{ mL} \times \frac{0.10 \text{ mol}}{1000 \text{ mL}} \times \frac{6.02214 \times 10^{23} \text{ molecules}}{1 \text{ mol}} = 6.0 \times 10^{20} \text{ molecules are}$$

in the first 10 mL aliquot of the solution. In order to find the number of times the volume must be doubled to get to one molecule, we can solve for the number of times this amount of molecules must be cut in half until it equals 1.

$$6.0 \times 10^{20} \text{ molecules} \times \left(\frac{1}{2}\right)^n = 1 \text{ molecule}$$

$$\log(6.0 \times 10^{20}) + n \log\left(\frac{1}{2}\right) = \log(1)$$

$$20.8 + n(-0.301) = 0$$

$$20.8 = 0.301n$$

$$n = 69$$

The other 21 additional doublings involve solutions with no X remaining. There can be no health benefits if there are no molecules of the active substance, X, left in the solution.

G.21 Find the volume of concentrated HCl solution that is equivalent to 10.0 L of the dilute HCl solution with respect to the number of moles of solute present in each volume.

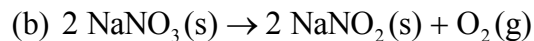
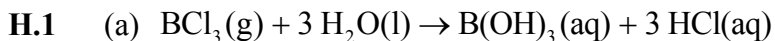
? mL con. HCl(aq)

$$= 10.0 \text{ L dilute HCl(aq)} \left(\frac{0.7436 \text{ mol HCl}}{1 \text{ L dilute HCl(aq)}} \right) \left(\frac{36.46 \text{ g HCl}}{1 \text{ mol HCl}} \right)$$

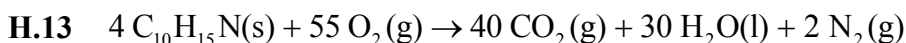
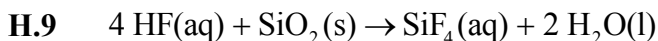
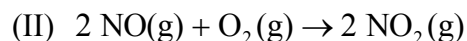
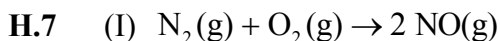
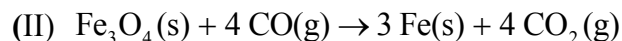
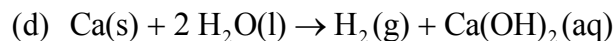
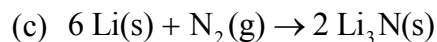
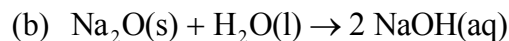
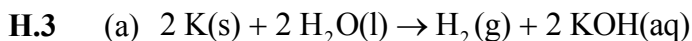
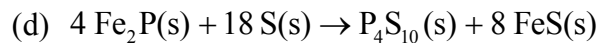
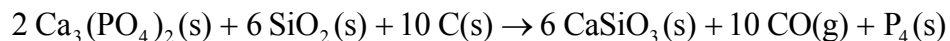
$$\times \left(\frac{100 \text{ g con. HCl(aq)}}{37.50 \text{ g HCl}} \right) \left(\frac{1 \text{ cm}^3}{1.205 \text{ g}} \right)$$

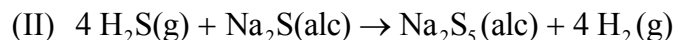
$$= 600. \text{ mL con. HCl}$$

Therefore 600. mL of concentrated HCl(aq) must be diluted up to a final volume of 10.0 L by adding water in order to form 0.7436 M HCl(aq).

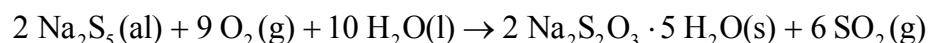


(c)





(III)



H.17 We can find the empirical formulas from the percent compositions.

First oxide:

$$\text{P} \quad 43.64 \text{ g} \div 30.97 \text{ g/mol} = 1.409 \text{ mol}$$

$$\text{O} \quad 56.36 \text{ g} \div 16.00 \text{ g/mol} = 3.523 \text{ mol}$$

$$\frac{3.523 \text{ mol}}{1.409 \text{ mol}} = 2.500 \quad 1:2.5 \text{ or } 2:5 \quad \text{P}_2\text{O}_5$$

Second oxide:

$$\text{P} \quad 56.34 \text{ g} \div 30.97 \text{ g/mol} = 1.819 \text{ mol}$$

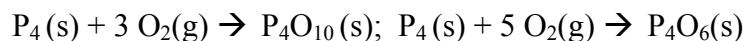
$$\text{O} \quad 43.66 \text{ g} \div 16.00 \text{ g/mol} = 2.729 \text{ mol}$$

$$\frac{2.729 \text{ mol}}{1.819 \text{ mol}} = 1.500 \quad 1:1.5 \text{ or } 2:3 \quad \text{P}_2\text{O}_3$$

These empirical formulas could be named diphosphorus pentoxide and diphosphorus trioxide. The names according to the Stock system are given below with the formulas of the actual compounds.

(a) P_4O_{10} (phosphorus (V) oxide), P_4O_6 (phosphorus (III) oxide);

(b) Since the molar masses of the empirical formulas are 142 g/mol and 110 g/mol respectively, and these masses are both half as big as the molar masses of the compounds, both molecular formulas are twice the empirical formulas.



I.1 (a) CH_3OH , nonelectrolyte; (b) CaBr_2 , strong electrolyte; (c) KI , strong electrolyte

I.3 (a) soluble; (b) slightly soluble; (c) insoluble; (d) insoluble

I.5 (a) $\text{Na}^+(\text{aq})$ and $\text{I}^-(\text{aq})$; (b)

$\text{Ag}^+(\text{aq})$ and $\text{CO}_3^{2-}(\text{aq})$, Ag_2CO_3 is insoluble. The very small amount that does go into solution will be present as Ag^+ and CO_3^{2-} ions.

(c) $\text{NH}_4^+(\text{aq})$ and $\text{PO}_4^{3-}(\text{aq})$ (d) $\text{Fe}^{2+}(\text{aq})$ and $\text{SO}_4^{2-}(\text{aq})$

I.7 (a) $\text{Fe}(\text{OH})_3$, precipitate; (b) Ag_2CO_3 , precipitate forms; (c) No precipitate will form because all possible products are soluble in water.

I.9 (a) net ionic equation: $\text{Fe}^{2+}(\text{aq}) + \text{S}^{2-}(\text{aq}) \rightarrow \text{FeS}(\text{s})$;

spectator ions: Na^+ , Cl^-

(b) net ionic equation: $\text{Pb}^{2+}(\text{aq}) + 2 \text{I}^-(\text{aq}) \rightarrow \text{PbI}_2(\text{s})$;

spectator ions: K^+ , NO_3^-

(c) net ionic equation: $\text{Ca}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{CaSO}_4(\text{s})$;

spectator ions: NO_3^- , K^+

(d) net ionic equation: $\text{Pb}^{2+}(\text{aq}) + \text{CrO}_4^{2-}(\text{aq}) \rightarrow \text{PbCrO}_4(\text{s})$;

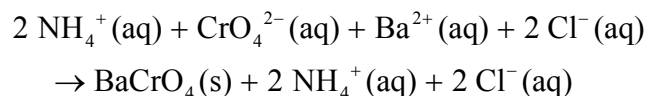
spectator ions: Na^+ , NO_3^-

(e) net ionic equation: $\text{Hg}_2^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{Hg}_2\text{SO}_4(\text{s})$;

spectator ions: K^+ , NO_3^-

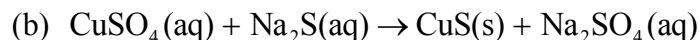
I.11 (a) overall equation: $(\text{NH}_4)_2\text{CrO}_4(\text{aq}) + \text{BaCl}_2(\text{aq})$
 $\rightarrow \text{BaCrO}_4(\text{s}) + 2 \text{NH}_4\text{Cl}(\text{aq})$

complete ionic equation:

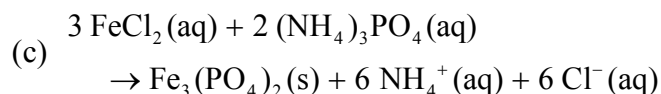
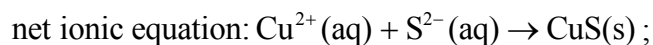
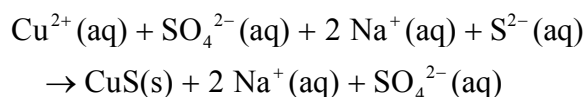


net ionic equation: $\text{Ba}^{2+}(\text{aq}) + \text{CrO}_4^{2-}(\text{aq}) \rightarrow \text{BaCrO}_4(\text{s})$;

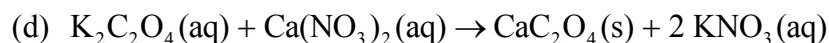
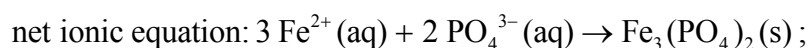
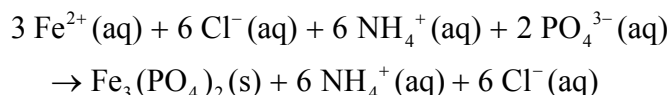
spectator ions: NH_4^+ , Cl^-



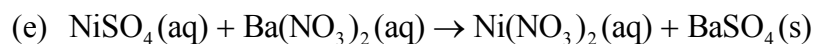
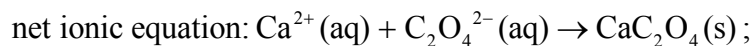
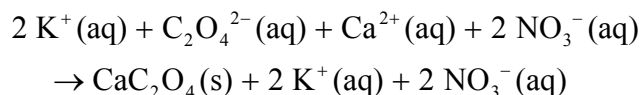
complete ionic equation:



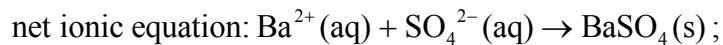
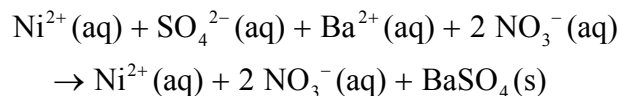
complete ionic equation:

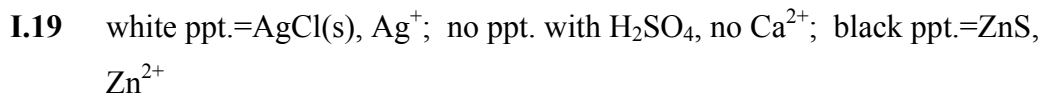
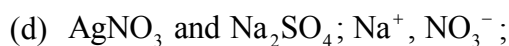
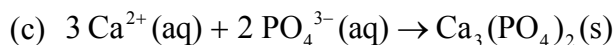
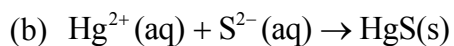
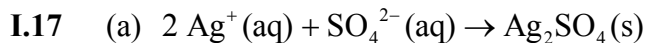
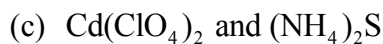
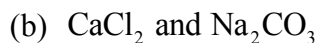
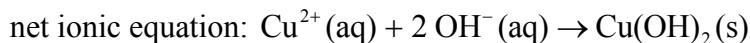
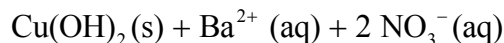
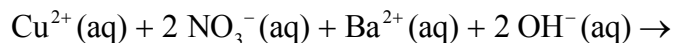
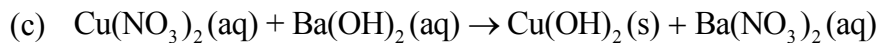
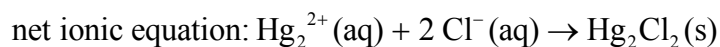
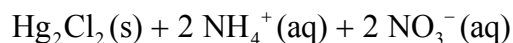
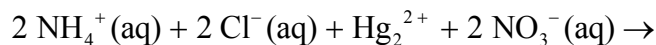
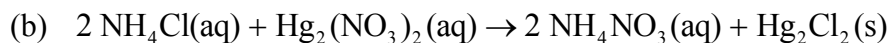
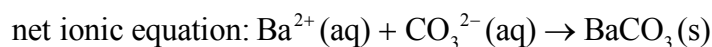
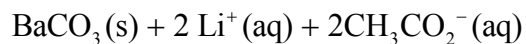
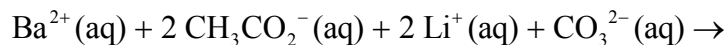
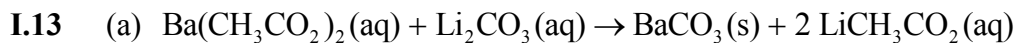


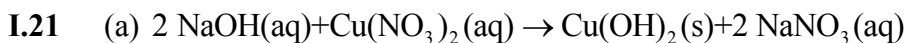
complete ionic equation:



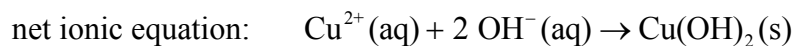
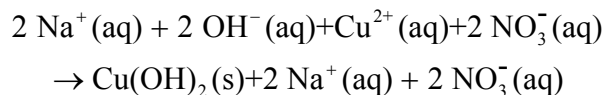
complete ionic equation:







complete ionic equation:



$$\frac{? \text{ mol Na}^+}{\text{L solution}} = \frac{20.00 \text{ mL NaOH(aq)} \left(\frac{0.100 \text{ mol Na}^+}{1000 \text{ mL NaOH(aq)}} \right)}{(20.00 + 40.00) \text{ mL solution}}$$

$$\begin{aligned} \text{(b)} \quad & \cdot \left(\frac{1000 \text{ mL}}{\text{L}} \right) \\ & = 0.0333 \text{ M Na}^+ \end{aligned}$$

I.23 (a) Find the number of moles of potassium chromate per liter of solution.

$$\begin{aligned} \frac{? \text{ mol K}_2\text{CrO}_4}{\text{L solution}} &= \left(\frac{3.50 \text{ g K}_2\text{CrO}_4}{75.0 \text{ mL solution}} \right) \left(\frac{1 \text{ mol K}_2\text{CrO}_4}{194.20 \text{ g K}_2\text{CrO}_4} \right) \left(\frac{1000 \text{ mL}}{1 \text{ L}} \right) \\ &= 0.240 \text{ M K}_2\text{CrO}_4(\text{aq}) \end{aligned}$$

(b) Find the number of grams of potassium in the amount of potassium chromate that was dissolved.

$$\begin{aligned} ? \text{ g K}^+ &= 3.50 \text{ g K}_2\text{CrO}_4 \left(\frac{1 \text{ mol K}_2\text{CrO}_4}{194.20 \text{ g K}_2\text{CrO}_4} \right) \left(\frac{2 \text{ mol K}^+}{1 \text{ mol K}_2\text{CrO}_4} \right) \\ &\quad \times \left(\frac{39.0983 \text{ g K}^+}{1 \text{ mol K}^+} \right) \\ &= 1.41 \text{ g K}^+ \end{aligned}$$

(c) MgCrO_4

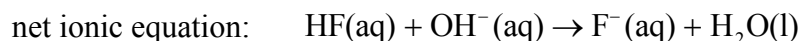
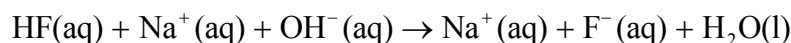


$$\begin{aligned}
 ? \text{ g Ag}_2\text{CrO}_4 &= 25.0 \text{ mL K}_2\text{CrO}_4(\text{aq}) \left(\frac{5.0 \text{ mol K}_2\text{CrO}_4}{1000 \text{ mL K}_2\text{CrO}_4(\text{aq})} \right) \\
 &\times \left(\frac{1 \text{ mol Ag}_2\text{CrO}_4}{1 \text{ mol K}_2\text{CrO}_4} \right) \left(\frac{331.74 \text{ g Ag}_2\text{CrO}_4}{1 \text{ mol Ag}_2\text{CrO}_4} \right) \\
 &= 41.5 \text{ g Ag}_2\text{CrO}_4 \text{ solid precipitate}
 \end{aligned}$$

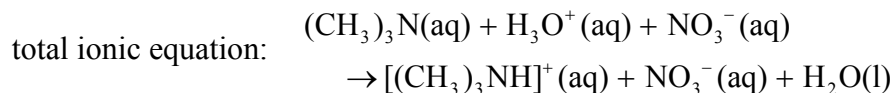
J.1 (a) base; (b) acid; (c) base; (d) acid; (e) base

J.3 (a) overall equation: $\text{HF}(\text{aq}) + \text{NaOH}(\text{aq}) \rightarrow \text{NaF}(\text{aq}) + \text{H}_2\text{O}(\text{l})$

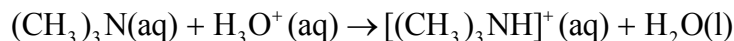
total ionic equation:



(b) overall equation: $(\text{CH}_3)_3\text{N}(\text{aq}) + \text{HNO}_3(\text{aq}) \rightarrow [(\text{CH}_3)_3\text{NH}]\text{NO}_3(\text{aq})$

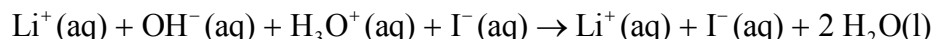


net ionic equation:



(c) overall equation: $\text{LiOH}(\text{aq}) + \text{HI}(\text{aq}) \rightarrow \text{LiI}(\text{aq}) + \text{H}_2\text{O}(\text{l})$

complete ionic equation:



J.5 (a) $\text{HBr}(\text{aq}) + \text{KOH}(\text{aq}) \rightarrow \text{KBr}(\text{aq}) + \text{H}_2\text{O}(\text{l})$

(b) $\text{Zn}(\text{OH})_2(\text{aq}) + 2 \text{HNO}_2(\text{aq}) \rightarrow \text{Zn}(\text{NO}_2)_2(\text{aq}) + 2 \text{H}_2\text{O}(\text{l})$

(c) $\text{Ca}(\text{OH})_2(\text{aq}) + 2 \text{HCN}(\text{aq}) \rightarrow \text{Ca}(\text{CN})_2(\text{aq}) + 2 \text{H}_2\text{O}(\text{l})$

(d) $3 \text{KOH}(\text{aq}) + \text{H}_3\text{PO}_4(\text{aq}) \rightarrow \text{K}_3\text{PO}_4(\text{aq}) + 3 \text{H}_2\text{O}(\text{l})$

J.7 (a) acid: $\text{H}_3\text{O}^+(\text{aq})$; base: $\text{CH}_3\text{NH}_2(\text{aq})$; (b) acid: $\text{HCl}(\text{aq})$;

base: $\text{C}_2\text{H}_5\text{NH}_2(\text{aq})$;

(c) acid: HI(aq); base: CaO(s)

J.9 Since X turns litmus red and conducts electricity poorly, it is a weak acid.

We can find the empirical formula from the percent composition.

$$\text{C } 26.68 \text{ g} \div 12.01 \text{ g/mol} = 2.221 \text{ mol}$$

$$\text{H } 2.239 \text{ g} \div 1.0079 \text{ g/mol} = 2.221 \text{ mol}$$

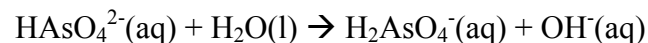
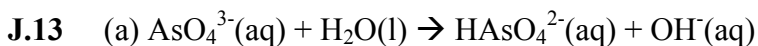
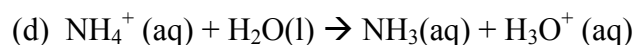
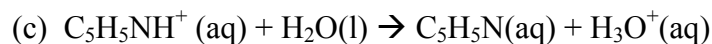
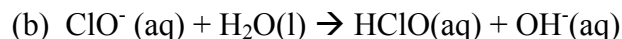
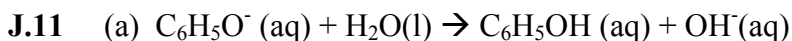
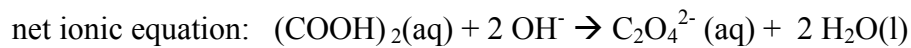
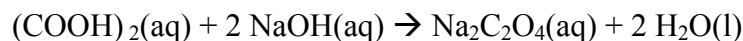
$$\text{O } 71.081 \text{ g} \div 16.00 \text{ g/mol} = 4.443 \text{ mol}$$

So the subscripts are 1:1:2 on the empirical formula.

(a) CHO₂;

(b) Since the molar mass of the empirical formula is 45.0 g/mol while the molar mass of X is 90.0 g/mol, the molecular formula is twice the empirical formula or C₂H₂O₄ ;.

(c) The weak acid whose formula matches the one given in part (b) is oxalic acid.



(b)

$$\begin{aligned} ? \text{ mol Na}^+ &= 35.0 \text{ g Na}_3\text{AsO}_4 \left(\frac{1 \text{ mol Na}_3\text{AsO}_4}{207.89 \text{ g Na}_3\text{AsO}_4} \right) \\ &\quad \times \left(\frac{3 \text{ mol Na}^+}{1 \text{ mol Na}_3\text{AsO}_4} \right) \\ &= 0.505 \text{ mol Na}^+ \end{aligned}$$

- K.1** (a) $2 \text{NO}_2(\text{g}) + \text{O}_3(\text{g}) \rightarrow \text{N}_2\text{O}_5(\text{g}) + \text{O}_2(\text{g})$
 (b) $\text{S}_8(\text{s}) + 16 \text{Na}(\text{s}) \rightarrow 8 \text{NaS}(\text{s})$
 (c) $2 \text{Cr}^{2+}(\text{aq}) + \text{Sn}^{4+}(\text{aq}) \rightarrow 2 \text{Cr}^{3+}(\text{aq}) + \text{Sn}^{2+}(\text{aq})$
 (d) $2 \text{As}(\text{s}) + 3 \text{Cl}_2(\text{g}) \rightarrow 2 \text{AsCl}_3(\text{l})$
- K.3** (a) $\text{Mg}^0(\text{s}) + \text{Cu}^{2+}(\text{aq}) \rightarrow \text{Mg}^{2+}(\text{aq}) + \text{Cu}^0(\text{s})$
 (b) $\text{Fe}^{2+}(\text{aq}) + \text{Ce}^{4+}(\text{aq}) \rightarrow \text{Fe}^{3+}(\text{aq}) + \text{Ce}^{3+}(\text{aq})$
 (c) $\text{H}_2(\text{g}) + \text{Cl}_2(\text{g}) \rightarrow 2 \text{HCl}(\text{g})$
 (d) $4 \text{Fe}(\text{s}) + 3 \text{O}_2(\text{g}) \rightarrow 2 \text{Fe}_2\text{O}_3(\text{s})$
- K.5** (a) +4; (b) +4; (c) +2; (d) +5; (e) +1; (f) 0
- K.7** (a) +2; (b) +2; (c) +6; (d) +4; (e) +1
- K.9** (a) Methanol $\text{CH}_3\text{OH}(\text{aq})$ is oxidized to formic acid (the carbon atom goes from an oxidation number of +2 to +4). The $\text{O}_2(\text{g})$ is reduced to O^{2-} present in water. (b) Mo is reduced from +5 to +4, while *some* sulfur (that which ends up as $\text{S}(\text{s})$) is oxidized from -2 to 0. The sulfur present in $\text{MoS}_2(\text{s})$ remains in the -2 oxidation state.
 (c) Tl^+ is both oxidized and reduced. The product $\text{Tl}(\text{s})$ is a reduction of Tl^+ (from +1 to 0) while the Tl^{3+} is produced via an oxidation of Tl^+ . A reaction in which a single substance is both oxidized and reduced is known as a *disproportionation reaction*.
- K.11** (a) Cl_2 will be reduced more easily and is therefore a stronger oxidizing agent than Cl^- .
 (b) N_2O_5 will be a stronger oxidizing agent because it will be readily reduced. N^{5+} will accept e^- more readily than will N^+ .

K.13 (a) oxidizing agent: H^+ in HCl(aq) ; reducing agent: Zn(s)

(b) oxidizing agent: $\text{SO}_2(\text{g})$; reducing agent: $\text{H}_2\text{S(g)}$

(c) oxidizing agent: $\text{B}_2\text{O}_3(\text{s})$; reducing agent: Mg(s)

K.15 (a) $\text{ClO}_3^- \rightarrow \text{ClO}_2$, Cl goes from +5 to +4; reducing agent

(b) $\text{SO}_4^{2-} \rightarrow \text{S}^{2-}$, S goes from +6 to -2; reducing agent

(c) $\text{Mn}^{2+} \rightarrow \text{MnO}_2$, Mn goes from +2 to +4; oxidizing agent

(d) $\text{HCHO} \rightarrow \text{HCOOH}$, C goes from 0 to +2; oxidizing agent

K.17 (a) oxidizing agent: $\text{WO}_3(\text{s})$; reducing agent: $\text{H}_2(\text{g})$

(b) oxidizing agent: HCl reducing agent: Mg(s)

(c) oxidizing agent: $\text{SnO}_2(\text{s})$; reducing agent: C(s)

(d) oxidizing agent: $\text{N}_2\text{O}_4(\text{g})$; reducing agent: $\text{N}_2\text{H}_4(\text{g})$

K.19 (a) $3 \text{N}_2\text{H}_4(\text{l}) \rightarrow 4 \text{NH}_3(\text{g}) + \text{N}_2(\text{g})$; (b) 2^- in N_2H_4 ; 3^- in NH_3 ; (c) N_2H_4 is both oxidizing and reducing agent;

(d) Factor label (dimensional analysis) can be used to find the volume of nitrogen.

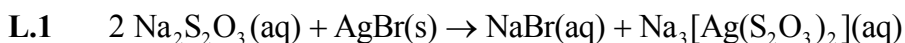
$$\begin{aligned} ? \text{ L N}_2(\text{g}) &= 1.0 \text{ L N}_2\text{H}_4(\text{l}) \left(\frac{1000 \text{ cm}^3}{1 \text{ L}} \right) \left(\frac{1.004 \text{ g}}{1 \text{ cm}^3} \right) \left(\frac{1 \text{ mol}}{32.0 \text{ g}} \right) \\ &\quad \times \left(\frac{1 \text{ mol N}_2}{3 \text{ mol N}_2\text{H}_4} \right) \left(\frac{28.0 \text{ g}}{1 \text{ mol}} \right) \left(\frac{24 \text{ L}}{28 \text{ g}} \right) \\ &= 2.5 \times 10^2 \text{ L N}_2(\text{g}) \end{aligned}$$

K.21 (a) $2 \text{Cr}^{2+} + \text{Cu}^{2+} \rightarrow 2 \text{Cr}^{3+} + \text{Cu(s)}$; (b) 2 e- transferred;

(c) Use factor label (dimensional analysis):

$$\begin{aligned} \frac{? \text{ mol NO}_3^-}{\text{L solution}} &= \left(\frac{50.5 \text{ g Cr(NO}_3)_2}{250.0 \text{ mL solution}} \right) \left(\frac{1 \text{ mol Cr(NO}_3)_2}{176.0 \text{ g Cr(NO}_3)_2} \right) \\ &\times \left(\frac{2 \text{ mol NO}_3^-}{1 \text{ mol Cr(NO}_3)_2} \right) \left(\frac{1000 \text{ mL}}{1 \text{ L}} \right) \\ &= 2.27 \text{ M NO}_3^- \end{aligned}$$

$$\begin{aligned} \frac{? \text{ mol SO}_4^{2-}}{\text{L solution}} &= \left(\frac{60.0 \text{ g CuSO}_4}{250.0 \text{ mL solution}} \right) \left(\frac{1 \text{ mol CuSO}_4}{159.62 \text{ g CuSO}_4} \right) \\ &\times \left(\frac{1 \text{ mol SO}_4^{2-}}{1 \text{ mol CuSO}_4} \right) \left(\frac{1000 \text{ mL}}{1 \text{ L}} \right) \\ &= 1.50 \text{ M SO}_4^{2-} \end{aligned}$$

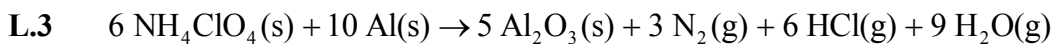


(a) moles of $\text{Na}_2\text{S}_2\text{O}_3$ needed to dissolve 1.0 mg AgBr

$$\begin{aligned} &= 1.0 \text{ mg AgBr} \left(\frac{1 \text{ g AgBr}}{1000 \text{ mg AgBr}} \right) \left(\frac{1 \text{ mol AgBr}}{187.78 \text{ g AgBr}} \right) \left(\frac{2 \text{ mol Na}_2\text{S}_2\text{O}_3}{1 \text{ mol AgBr}} \right) \\ &= 1.1 \times 10^{-5} \text{ mol Na}_2\text{S}_2\text{O}_3 \end{aligned}$$

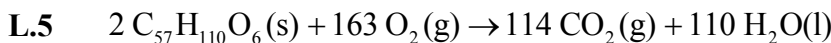
(b) mass of AgBr to produce 0.033 mol $\text{Na}_3[\text{Ag}(\text{S}_2\text{O}_3)_2]$

$$\begin{aligned} &= 0.033 \text{ mol Na}_3[\text{Ag}(\text{S}_2\text{O}_3)_2] \left(\frac{1 \text{ mol AgBr}}{1 \text{ mol Na}_3[\text{Ag}(\text{S}_2\text{O}_3)_2]} \right) \left(\frac{187.78 \text{ g AgBr}}{1 \text{ mol AgBr}} \right) \\ &= 6.2 \text{ g AgBr} \end{aligned}$$



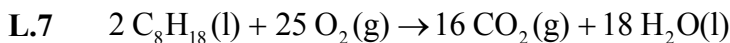
$$\begin{aligned} \text{(a)} \quad &(1.325 \text{ kg NH}_4\text{ClO}_4) \left(\frac{1 \text{ mol NH}_4\text{ClO}_4}{117.49 \text{ g NH}_4\text{ClO}_4} \right) \left(\frac{1000 \text{ g}}{1 \text{ kg}} \right) \\ &\left(\frac{10 \text{ mol Al}}{6 \text{ mol NH}_4\text{ClO}_4} \right) \left(\frac{26.98 \text{ g Al}}{1 \text{ mol Al}} \right) = 507.1 \text{ g Al} \end{aligned}$$

$$\begin{aligned}
 \text{(b)} \quad & 3500 \text{ kg Al} \left(\frac{1000 \text{ g Al}}{1 \text{ kg Al}} \right) \left(\frac{1 \text{ mol Al}}{26.98 \text{ g Al}} \right) \\
 & \times \left(\frac{5 \text{ mol Al}_2\text{O}_3}{10 \text{ mol Al}} \right) \left(\frac{101.96 \text{ g Al}_2\text{O}_3}{1 \text{ mol Al}_2\text{O}_3} \right) \\
 & = 6.613 \times 10^6 \text{ g Al}_2\text{O}_3 \text{ or } 6.613 \times 10^3 \text{ kg Al}_2\text{O}_3
 \end{aligned}$$



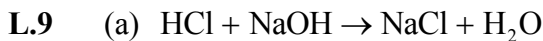
$$\begin{aligned}
 \text{(a)} \quad & (454 \text{ g fat}) \left(\frac{1 \text{ mol fat}}{891.44 \text{ g fat}} \right) \left(\frac{110 \text{ mol H}_2\text{O}}{2 \text{ mol fat}} \right) \left(\frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} \right) \\
 & = 505 \text{ g H}_2\text{O}
 \end{aligned}$$

$$\begin{aligned}
 \text{(b)} \quad & (454 \text{ g fat}) \left(\frac{1 \text{ mol fat}}{891.44 \text{ g}} \right) \left(\frac{163 \text{ mol O}_2}{2 \text{ mol fat}} \right) \left(\frac{32.00 \text{ g O}_2}{1 \text{ mol O}_2} \right) \\
 & = 1.33 \times 10^3 \text{ g O}_2
 \end{aligned}$$



$d = 0.79 \text{ g} \cdot \text{mL}^{-1}$, density of gasoline

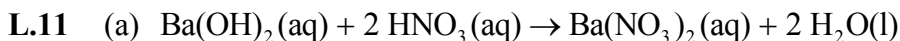
$$\begin{aligned}
 & (3.785 \text{ L gas}) \left(\frac{1000 \text{ mL}}{1 \text{ L}} \right) \left(\frac{0.79 \text{ g gas}}{1 \text{ mL}} \right) \left(\frac{1 \text{ mol gas}}{114.22 \text{ g gas}} \right) \left(\frac{18 \text{ mol H}_2\text{O}}{2 \text{ mol C}_8\text{H}_{18}} \right) \\
 & \left(\frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} \right) = 4.246 \times 10^3 \text{ g H}_2\text{O} \text{ or } 4.2 \text{ kg H}_2\text{O}
 \end{aligned}$$



$$17.40 \text{ mL} \left(\frac{0.234 \text{ mol HCl}}{1000 \text{ mL}} \right) \left(\frac{1 \text{ mol NaOH}}{1 \text{ mol HCl}} \right) = 0.00407 \text{ mol}$$

$$\text{concentration of NaOH} = \frac{0.00407 \text{ mol}}{15.00 \times 10^{-3} \text{ L}} = 0.271 \text{ M}$$

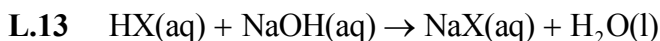
$$\text{(b)} \quad (0.271 \text{ mol} \cdot \text{L}^{-1})(0.01500 \text{ L})(40.00 \text{ g} \cdot \text{mol}^{-1}) = 0.163 \text{ g NaOH}$$



$$\begin{aligned}
 \frac{? \text{ mol HNO}_3}{\text{L HNO}_3(\text{aq})} &= (11.56 \text{ mL Ba(OH)}_2(\text{aq})) \left(\frac{9.670 \text{ g Ba(OH)}_2}{250. \text{ mL Ba(OH)}_2(\text{aq})} \right) \\
 &\quad \times \frac{\left(\frac{1 \text{ mol Ba(OH)}_2}{171.36 \text{ g Ba(OH)}_2} \right) \left(\frac{2 \text{ mol HNO}_3}{1 \text{ mol Ba(OH)}_2} \right)}{(25.0 \text{ mL HNO}_3(\text{aq})) \left(\frac{1 \text{ L}}{1000 \text{ mL}} \right)} \\
 &= 0.209 \text{ mol} \cdot \text{L}^{-1}
 \end{aligned}$$

(b) mass of HNO_3 in solution:

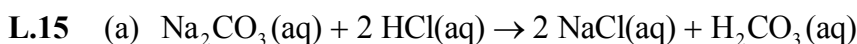
$$\left(\frac{0.209 \text{ mol}}{1000 \text{ mL}} \right) (25.0 \text{ mL}) \left(\frac{63.02 \text{ g HNO}_3}{1 \text{ mol HNO}_3} \right) = 0.329 \text{ g}$$



$$(68.8 \text{ mL}) \left(\frac{0.750 \text{ mol NaOH}}{1000 \text{ mL NaOH}} \right) = 0.0516 \text{ mol NaOH}$$

3.25 g HX corresponds to 0.0516 mol NaOH used

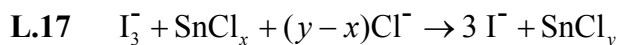
$$\frac{3.25 \text{ g}}{0.0516 \text{ mol}} = 63.0 \text{ g} \cdot \text{mol}^{-1} = \text{molar mass of acid}$$



(b) First find the concentration of the diluted acid.

$$\begin{aligned}
 ? \text{ M HCl}(\text{aq}) \text{ dilute} &= \left(\frac{0.832 \text{ g Na}_2\text{CO}_3}{0.100 \text{ L base solution}} \right) \left(\frac{1 \text{ mol Na}_2\text{CO}_3}{105.99 \text{ g Na}_2\text{CO}_3} \right) \\
 &\quad \times \left(\frac{0.025 \text{ L base solution}}{0.03125 \text{ L acid solution}} \right) \\
 &\quad \times \left(\frac{2 \text{ mol HCl}}{1 \text{ mol Na}_2\text{CO}_3} \right) = 0.126 \text{ M HCl}(\text{aq}) \text{ dilute}
 \end{aligned}$$

The original HCl solution is 100 times more concentrated than the solution used for titration (diluted 10.00 mL to 1000 mL), so the original concentration of the HCl solution is $12.6 \text{ mol} \cdot \text{L}^{-1}$.



The information given can be used to find the molar mass of the reactant in order to identify it.

$$25.00 \text{ mL} \left(\frac{0.120 \text{ mol I}_3^-}{1000 \text{ mL}} \right) = 3.00 \times 10^{-3} \text{ mol I}_3^-$$

$$30.00 \text{ mL} \left(\frac{19.0 \text{ g tin chloride}}{1000 \text{ mL}} \right) = 0.570 \text{ g tin chloride}$$

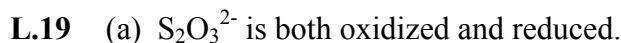
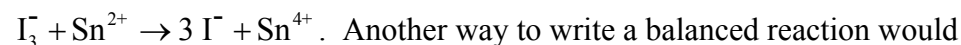
If the reaction is 1:1 then the # moles of I_3^- is the same as the number of moles of SnCl_x . In that case, the molar mass of the tin chloride reactant is

$$\frac{0.570 \text{ g}}{3.00 \times 10^{-3} \text{ mol}} = 190. \text{ g} \cdot \text{mol}^{-1}. \text{ This molar mass matches that of } \text{SnCl}_2,$$

$189.61 \text{ g} \cdot \text{mol}^{-1}$. Tin (II) chloride also has the correct mass percent tin.

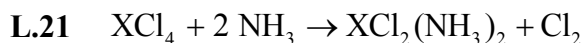
$$\frac{118.71 \text{ g} \cdot \text{mol}^{-1} \text{ Sn}}{189.61 \text{ g} \cdot \text{mol}^{-1} \text{ SnCl}_2} \times 100 = 62.6\%$$

Since the product compound is oxidized relative to the reactant, we can expect it to be Sn(IV). The net ionic equation for the reaction is



(b) Find the number of grams of thiosulfate ion in 10.1 mL of solution.

$$\begin{aligned} ? \text{ g S}_2\text{O}_3^{2-} &= 10.1 \text{ mL HSO}_3^-(\text{aq}) \left(\frac{1.45 \text{ g HSO}_3^-(\text{aq})}{1 \text{ mL HSO}_3^-(\text{aq})} \right) \left(\frac{55.0 \text{ g HSO}_3^-}{100 \text{ g HSO}_3^-(\text{aq})} \right) \\ &\quad \times \left(\frac{1 \text{ mol HSO}_3^-}{81.0 \text{ g HSO}_3^-} \right) \left(\frac{1 \text{ mol S}_2\text{O}_3^{2-}}{1 \text{ mol HSO}_3^-} \right) \left(\frac{112.0 \text{ g S}_2\text{O}_3^{2-}}{1 \text{ mol S}_2\text{O}_3^{2-}} \right) \\ &= 11.1 \text{ g S}_2\text{O}_3^{2-} \text{ present initially} \end{aligned}$$



The reactant and product that contain X are in a 1:1 ratio, so 3.571 g of the reactant is equivalent to 3.180 g of the product. The molar mass of the reactant is $x + 4(35.453 \text{ g/mol})$ while that of the product is $x + 2(35.453 \text{ g/mol}) + 2(14.01 \text{ g/mol}) + 6(1.0079 \text{ g/mol})$. Therefore we can set up the following proportion in order to solve for x in g/mol:

$$\frac{x + 141.8}{x + 104.97} = \frac{3.571}{3.180}$$

$$1.1230x - x = 23.923$$

$$x = 194.6 \text{ g} \cdot \text{mol}^{-1}, \text{ or Pt}$$

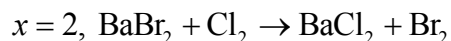
L.23 The number of moles of product is

$$2.27 \text{ g} \div 208.23 \text{ g} \cdot \text{mol}^{-1} = 0.0109 \text{ moles BaCl}_2. \text{ An equivalent number}$$

of moles is represented by 3.25 g of BaBr_x , so its molar mass is

$$3.25 \text{ g} \div 0.0109 \text{ moles} = 298 \text{ g} \cdot \text{mol}^{-1} \text{ BaBr}_x. \text{ Since } 137.33 \text{ g is}$$

attributable to Ba, 161 g must be Br. Each Br has a mass of 80.4 g/mol so there must be 2 moles of Br for each mole of Ba in the reactant.



M.1 $\text{CaCO}_3(\text{s}) \rightarrow \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$

theoretical yield:

$$(42.73 \text{ g CaCO}_3) \left(\frac{1 \text{ mol CaCO}_3}{100.09 \text{ g CaCO}_3} \right) \left(\frac{1 \text{ mol CO}_2}{1 \text{ mol CaCO}_3} \right) \left(\frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2} \right) \\ = 18.79 \text{ g CO}_2$$

actual yield:

$$\frac{17.5 \text{ g}}{18.79 \text{ g}} \times 100\% = 93.1\% \text{ yield}$$

M.3 $\text{C}_x\text{H}_y\text{Cl}_z + (x + \frac{y}{2})\text{O}_2 \rightarrow x \text{ CO}_2 + y \text{ H}_2\text{O} + \frac{z}{2}\text{Cl}_2$

$$\frac{1.52 \text{ g}}{360.88 \text{ g} \cdot \text{mol}^{-1}} = 4.21 \times 10^{-3} \text{ mol Arochlor yields } \frac{2.224 \text{ g}}{44.0 \text{ g} \cdot \text{mol}^{-1}} \\ = 5.055 \times 10^{-2} \text{ mol CO}_2$$

$$\frac{2.53 \text{ g}}{360.88 \text{ g} \cdot \text{mol}^{-1}} = 7.01 \times 10^{-3} \text{ mol Arochlor yields } \frac{0.2530 \text{ g}}{18.01 \text{ g} \cdot \text{mol}^{-1}} \\ = 1.405 \times 10^{-2} \text{ mol H}_2\text{O}$$

$$\text{Therefore, } x = \frac{5.055 \times 10^{-2}}{4.21 \times 10^{-3}} = 12.0 \text{ and } y = \frac{1.405 \times 10^{-2}}{7.01 \times 10^{-3}} = 2.00$$

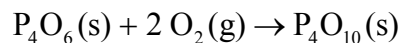
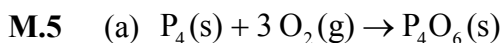
$$12.011x + 1.0079y + 35.453z = 360.88$$

$$12.011(12.0) + 1.0079(2.00) + 35.453z = 360.88$$

$$35.453z = 214.7$$

$$z = 6.06$$

Since the number of Cl atoms per Arochlor 1254 molecule must be a whole number, the number of chlorine atoms is 6.



In the first reaction, 5.77 g P_4 uses

$$(5.77 \text{ g P}_4) \left(\frac{1 \text{ mol P}_4}{123.88 \text{ g P}_4} \right) \left(\frac{3 \text{ mol O}_2}{1 \text{ mol P}_4} \right) \left(\frac{32.00 \text{ g O}_2}{1 \text{ mol O}_2} \right) = 4.47 \text{ g O}_2 (\text{g})$$

$$\text{excess O}_2 = 5.77 \text{ g} - 4.47 \text{ g O}_2 = 1.30 \text{ g O}_2$$

In the second reaction, 5.77 g P_4 uses

$$\left(\frac{5.77 \text{ g P}_4}{123.88 \text{ g} \cdot \text{mol}^{-1} \text{ P}_4} \right) \left(\frac{1 \text{ mol P}_4\text{O}_6}{1 \text{ mol P}_4} \right) \left(\frac{2 \text{ mol O}_2}{1 \text{ mol P}_4\text{O}_6} \right) \left(\frac{32.00 \text{ g O}_2}{1 \text{ mol O}_2} \right) \\ = 2.98 \text{ g O}_2$$

limiting reagent: O_2

$$(b) \left(\frac{1.30 \text{ g O}_2}{32.00 \text{ g} \cdot \text{mol}^{-1} \text{ O}_2} \right) \left(\frac{1 \text{ mol P}_4\text{O}_{10}}{2 \text{ mol O}_2} \right) \left(\frac{283.88 \text{ g P}_4\text{O}_{10}}{1 \text{ mol P}_4\text{O}_{10}} \right) = 5.77 \text{ g P}_4\text{O}_{10}$$

$$\begin{aligned} \text{(c)} \quad & \left(\frac{1.30 \text{ g O}_2}{32.00 \text{ g} \cdot \text{mol}^{-1} \text{ O}_2} \right) \left(\frac{1 \text{ mol P}_4\text{O}_6}{2 \text{ mol O}_2} \right) \left(\frac{219.88 \text{ g P}_4\text{O}_6}{1 \text{ mol P}_4\text{O}_6} \right) \\ & = 4.47 \text{ g P}_4\text{O}_6 \text{ used} \end{aligned}$$

In the first reaction, 5.77 g P₄ produces

$$\left(\frac{5.77 \text{ g P}_4}{123.88 \text{ g} \cdot \text{mol}^{-1}} \right) \left(\frac{219.88 \text{ g P}_4\text{O}_6}{1 \text{ mol P}_4\text{O}_6} \right) \left(\frac{1 \text{ mol P}_4\text{O}_6}{1 \text{ mol P}_4} \right) = 10.2 \text{ g P}_4\text{O}_6$$

excess reagent: 10.2 g – 4.47 g = 5.7 g P₄O₆

M.7 C₆₃H₈₈CoN₁₄O₁₄P. The molar mass of cobalamin is 1355.37 g · mol⁻¹.

$$n_{\text{cobalamin}} = \frac{0.1674 \text{ g}}{1355.37 \text{ g} \cdot \text{mol}^{-1}} = 1.235 \times 10^{-4} \text{ mol}$$

1 mole of cobalamin will produce 63 moles of CO₂ and 44 moles of H₂O.

$$\begin{aligned} m_{\text{CO}_2} &= (1.235 \times 10^{-4} \text{ mol cobalamin}) \left(\frac{63 \text{ mol CO}_2}{1 \text{ mol cobalamin}} \right) \\ &\quad \times (44.01 \text{ g} \cdot \text{mol}^{-1} \text{ CO}_2) \\ &= 0.3424 \text{ g CO}_2 \end{aligned}$$

$$\begin{aligned} m_{\text{H}_2\text{O}} &= (1.235 \times 10^{-4} \text{ mol cobalamin}) \left(\frac{44 \text{ mol H}_2\text{O}}{1 \text{ mol cobalamin}} \right) \\ &\quad \times (18.02 \text{ g} \cdot \text{mol}^{-1} \text{ H}_2\text{O}) \\ &= 0.09792 \text{ g H}_2\text{O} \end{aligned}$$

M.9 $(0.682 \text{ g CO}_2) \left(\frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} \right) \left(\frac{1 \text{ mol C}}{1 \text{ mol CO}_2} \right) = 0.0155 \text{ mol C}$

$$(0.0155 \text{ mol C})(12.01 \text{ g} \cdot \text{mol}^{-1} \text{ C}) = 0.186 \text{ g C}$$

$$(0.174 \text{ g H}_2\text{O}) \left(\frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \right) \left(\frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} \right) = 0.0193 \text{ mol H}$$

$$(0.0193 \text{ mol H})(1.0079 \text{ g} \cdot \text{mol}^{-1} \text{ H}) = 0.0195 \text{ g H}$$

$$(0.110 \text{ g N}_2) \left(\frac{1 \text{ mol N}_2}{28.02 \text{ g N}_2} \right) \left(\frac{2 \text{ mol N}}{1 \text{ mol N}_2} \right) = 0.00785 \text{ mol N}$$

$$(0.00785 \text{ mol N})(14.01 \text{ g} \cdot \text{mol}^{-1} \text{ N}) = 0.110 \text{ g N}$$

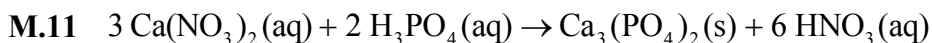
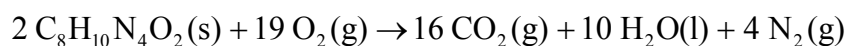
$$\text{mass of O} = 0.376 \text{ g} - (0.186 \text{ g} + 0.0193 \text{ g} + 0.110 \text{ g}) = 0.061 \text{ g O}$$

$$\frac{0.061 \text{ g O}}{16.00 \text{ g O}} = 0.0038 \text{ mol O}$$

Dividing each amount by 0.0038 gives C : H : N : O ratios = 4.1 : 5.1 : 2.1 : 1. The empirical formula is $\text{C}_4\text{H}_5\text{N}_2\text{O}$.

The molecular mass of caffeine is $194 \text{ g} \cdot \text{mol}^{-1}$. Its empirical mass is $97.10 \text{ g} \cdot \text{mol}^{-1}$.

$$\text{molecular formula} = 2 \times \text{empirical formula} = \text{C}_8\text{H}_{10}\text{N}_4\text{O}_2$$



(a) The solid is calcium phosphate, $\text{Ca}_3(\text{PO}_4)_2$.

$$\begin{aligned} \text{(b)} \quad (206 \text{ g Ca}(\text{NO}_3)_2) & \left(\frac{1 \text{ mol Ca}(\text{NO}_3)_2}{164.10 \text{ g Ca}(\text{NO}_3)_2} \right) \left(\frac{2 \text{ mol H}_3\text{PO}_4}{3 \text{ mol Ca}(\text{NO}_3)_2} \right) \\ & \left(\frac{97.99 \text{ g H}_3\text{PO}_4}{1 \text{ mol H}_3\text{PO}_4} \right) = 82.01 \text{ g H}_3\text{PO}_4 \end{aligned}$$

Therefore $\text{Ca}(\text{NO}_3)_2$ is the limiting reagent.

$$\begin{aligned} (206 \text{ g Ca}(\text{NO}_3)_2) & \left(\frac{1 \text{ mol Ca}(\text{NO}_3)_2}{164.10 \text{ g Ca}(\text{NO}_3)_2} \right) \left(\frac{1 \text{ mol Ca}_3(\text{PO}_4)_2}{3 \text{ mol Ca}(\text{NO}_3)_2} \right) \\ & \left(\frac{310.18 \text{ g Ca}_3(\text{PO}_4)_2}{1 \text{ mol Ca}_3(\text{PO}_4)_2} \right) = 130. \text{ g Ca}_3(\text{PO}_4)_2 \end{aligned}$$

M.13 If the 2-naphthol ($144.16 \text{ g} \cdot \text{mol}^{-1}$) were pure, it would give the following combustion analysis:

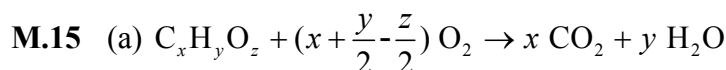
$$\%C = \frac{10(12.01 \text{ g} \cdot \text{mol}^{-1})}{(144.16 \text{ g} \cdot \text{mol}^{-1} \text{ naphthol})} \times 100\% = 83.31\% \text{ C}$$

$$\%H = \frac{8(1.0079 \text{ g} \cdot \text{mol}^{-1})}{(144.16 \text{ g} \cdot \text{mol}^{-1} \text{ naphthol})} \times 100\% = 5.59\% \text{ H}$$

The observed percentages are low as is expected for a sample contaminated with a substance that contains no C or H. Because the sample does not contain C or H, the percent purity can be easily obtained by

$$\% \text{purity (based on C)} = \frac{\% \text{ found}}{\% \text{ theoretical}} = \frac{77.48\% \text{ mixture}}{83.31\% \text{ pure naphthol}} \times 100\% = 93.00\%$$

$$\% \text{purity (based on H)} = \frac{\% \text{ found}}{\% \text{ theoretical}} = \frac{5.20\% \text{ mixture}}{5.59\% \text{ pure naphthol}} \times 100\% = 93.0\%$$



$$\frac{2.492 \text{ g CO}_2}{44.0 \text{ g} \cdot \text{mol}^{-1}} = 5.664 \times 10^{-2} \text{ mol CO}_2 = \text{mol C}$$

$$\frac{0.6495 \text{ g H}_2\text{O}}{18.01 \text{ g} \cdot \text{mol}^{-1}} = 3.608 \times 10^{-2} \text{ mol H}_2\text{O} = 7.216 \times 10^{-2} \text{ mol H}$$

$$(5.664 \times 10^{-2} \text{ mol C}) \left(\frac{12.011 \text{ g}}{\text{mol}} \right) = 0.6803 \text{ g C}$$

$$(7.216 \times 10^{-2} \text{ mol H}) \left(\frac{1.0079 \text{ g}}{\text{mol}} \right) = 0.07273 \text{ g H}$$

$$1.000 \text{ g compound} - (0.6803 \text{ g} + 0.07273 \text{ g}) = 0.2470 \text{ g}$$

$$0.2470 \text{ g O} \div 16.00 \text{ g} \cdot \text{mol}^{-1} = 1.544 \times 10^{-2} \text{ mol O}$$

$$\frac{5.664 \times 10^{-2}}{1.544 \times 10^{-2}} = 3.67 \quad \frac{7.216 \times 10^{-2}}{1.544 \times 10^{-2}} = 4.67$$

The mole ratio of C:H:O is 3.67:4.67:1 or 11:14:3, so the empirical formula is $\text{C}_{11}\text{H}_{14}\text{O}_3$.

(b) The molar mass of the empirical formula is 194 g/mol, which is half of 388.46 g/mol. Therefore, the molecular formula of the compound is $\text{C}_{22}\text{H}_{28}\text{O}_6$.

M.17 Determine the number of moles of each element present in the compound then find their ratios to get the subscripts for the empirical formula.

$$\frac{0.055 \text{ g Cl}}{35.453 \text{ g} \cdot \text{mol}^{-1}} = 1.55 \times 10^{-3} \text{ mol Cl}$$

$$\frac{0.0682 \text{ g CO}_2}{44.0 \text{ g} \cdot \text{mol}^{-1}} = 1.55 \times 10^{-3} \text{ mol CO}_2 = \text{mol C}$$

$$\frac{0.0140 \text{ g H}_2\text{O}}{18.01 \text{ g} \cdot \text{mol}^{-1}} = 7.78 \times 10^{-4} \text{ mol H}_2\text{O} = 1.56 \times 10^{-3} \text{ mol H}$$

0.100 g compound

$$\begin{aligned} & - \left(0.055 \text{ g Cl} + 1.55 \times 10^{-3} \text{ mol} \cdot \frac{12.0 \text{ g C}}{\text{mol}} + 1.55 \times 10^{-3} \text{ mol} \cdot \frac{1.0079 \text{ g H}}{\text{mol}} \right) \\ & = 0.0247 \text{ g O} \\ & \times \left(\frac{1 \text{ mol}}{16.00 \text{ g}} \right) = 1.55 \times 10^{-3} \text{ mol O} \end{aligned}$$

So the mole ratio is 1:1:1:1 and the empirical formula is CHOC1.

There is not sufficient information to allow the determination of the molecular formula.