

CHAPTER 4

TYPES OF CHEMICAL REACTIONS AND SOLUTION STOICHIOMETRY

Aqueous Solutions: Strong and Weak Electrolytes

10. Only statement b is true. A concentrated solution can also contain a nonelectrolyte dissolved in water, for example, concentrated sugar water. Acids are either strong or weak electrolytes. Some ionic compounds are not soluble in water, so they are not labeled as a specific type of electrolyte.
11. a. Polarity is a term applied to covalent compounds. Polar covalent compounds have an unequal sharing of electrons in bonds that results in unequal charge distribution in the overall molecule. Polar molecules have a partial negative end and a partial positive end. These are not full charges as in ionic compounds but are charges much smaller in magnitude. Water is a polar molecule and dissolves other polar solutes readily. The oxygen end of water (the partial negative end of the polar water molecule) aligns with the partial positive end of the polar solute, whereas the hydrogens of water (the partial positive end of the polar water molecule) align with the partial negative end of the solute. These opposite charge attractions stabilize polar solutes in water. This process is called hydration. Nonpolar solutes do not have permanent partial negative and partial positive ends; nonpolar solutes are not stabilized in water and do not dissolve.
- b. KF is a soluble ionic compound, so it is a strong electrolyte. KF(aq) actually exists as separate hydrated K^+ ions and hydrated F^- ions in solution. $C_6H_{12}O_6$ is a polar covalent molecule that is a nonelectrolyte. $C_6H_{12}O_6$ is hydrated as described in part a.
- c. RbCl is a soluble ionic compound, so it exists as separate hydrated Rb^+ ions and hydrated Cl^- ions in solution. AgCl is an insoluble ionic compound so the ions stay together in solution and fall to the bottom of the container as a precipitate.
- d. HNO_3 is a strong acid and exists as separate hydrated H^+ ions and hydrated NO_3^- ions in solution. CO is a polar covalent molecule and is hydrated as explained in part a.
12. The electrolyte designation refers to how completely the dissolved solute breaks up into ions. Strong electrolytes completely break up into ions when in water, weak electrolytes only partially break up into ions (less than 5% usually), and nonelectrolytes do not break up into ions when they dissolve in water. The conductivity apparatus illustrated in Figure 4.4 of the text is one way to experimentally determine the type of electrolyte. As illustrated, a bright light indicates many charge carriers (ions) are present and the solute is a strong electrolyte. A dim light indicates few ions are present so the solute is a weak electrolyte, and no light indicates no ions are present so the solute is a nonelectrolyte.

13. a. $\text{Ba}(\text{NO}_3)_2(\text{aq}) \rightarrow \text{Ba}^{2+}(\text{aq}) + 2 \text{NO}_3^{-}(\text{aq})$; picture iv represents the Ba^{2+} and NO_3^{-} ions present in $\text{Ba}(\text{NO}_3)_2(\text{aq})$.
- b. $\text{NaCl}(\text{aq}) \rightarrow \text{Na}^{+}(\text{aq}) + \text{Cl}^{-}(\text{aq})$; picture ii represents $\text{NaCl}(\text{aq})$.
- c. $\text{K}_2\text{CO}_3(\text{aq}) \rightarrow 2 \text{K}^{+}(\text{aq}) + \text{CO}_3^{2-}(\text{aq})$; picture iii represents $\text{K}_2\text{CO}_3(\text{aq})$.
- d. $\text{MgSO}_4(\text{aq}) \rightarrow \text{Mg}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq})$; picture i represents $\text{MgSO}_4(\text{aq})$.
14. $\text{MgSO}_4(\text{s}) \rightarrow \text{Mg}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq})$; $\text{NH}_4\text{NO}_3(\text{s}) \rightarrow \text{NH}_4^{+}(\text{aq}) + \text{NO}_3^{-}(\text{aq})$
15. Solution A: $\frac{4 \text{ molecules}}{1.0 \text{ L}}$; solution B: $\frac{6 \text{ molecules}}{4.0 \text{ L}} = \frac{1.5 \text{ molecules}}{1.0 \text{ L}}$
- Solution C: $\frac{4 \text{ molecules}}{2.0 \text{ L}} = \frac{2 \text{ molecules}}{1.0 \text{ L}}$; solution D: $\frac{6 \text{ molecules}}{2.0 \text{ L}} = \frac{3 \text{ molecules}}{1.0 \text{ L}}$

Solution A has the most molecules per unit volume so solution A is most concentrated. This is followed by solution D, then solution C. Solution B has the fewest molecules per unit volume, so solution B is least concentrated.

Solution Concentration: Molarity

16. a. $1.00 \text{ L solution} \times \frac{0.50 \text{ mol H}_2\text{SO}_4}{\text{L}} = 0.50 \text{ mol H}_2\text{SO}_4$
- $0.50 \text{ mol H}_2\text{SO}_4 \times \frac{1 \text{ L}}{18 \text{ mol H}_2\text{SO}_4} = 2.8 \times 10^{-2} \text{ L concentrated H}_2\text{SO}_4 \text{ or } 28 \text{ mL}$

Dilute 28 mL of concentrated H_2SO_4 to a total volume of 1.00 L with water. The resulting 1.00 L of solution will be a 0.50 M H_2SO_4 solution.

- b. We will need 0.50 mol HCl.

$$0.50 \text{ mol HCl} \times \frac{1 \text{ L}}{12 \text{ mol HCl}} = 4.2 \times 10^{-2} \text{ L} = 42 \text{ mL}$$

Dilute 42 mL of concentrated HCl to a final volume of 1.00 L.

- c. We need 0.50 mol NiCl_2 .

$$0.50 \text{ mol NiCl}_2 \times \frac{1 \text{ mol NiCl}_2 \cdot 6\text{H}_2\text{O}}{\text{mol NiCl}_2} \times \frac{237.69 \text{ g NiCl}_2 \cdot 6\text{H}_2\text{O}}{\text{mol NiCl}_2 \cdot 6\text{H}_2\text{O}} = 118.8 \text{ g NiCl}_2 \cdot 6\text{H}_2\text{O} \approx 120 \text{ g}$$

Dissolve 120 g $\text{NiCl}_2 \cdot 6\text{H}_2\text{O}$ in water, and add water until the total volume of the solution is 1.00 L.

$$d. \quad 1.00 \text{ L} \times \frac{0.50 \text{ mol HNO}_3}{\text{L}} = 0.50 \text{ mol HNO}_3$$

$$0.50 \text{ mol HNO}_3 \times \frac{1 \text{ L}}{16 \text{ mol HNO}_3} = 0.031 \text{ L} = 31 \text{ mL}$$

Dissolve 31 mL of concentrated reagent in water. Dilute to a total volume of 1.00 L.

e. We need 0.50 mol Na_2CO_3 .

$$0.50 \text{ mol Na}_2\text{CO}_3 \times \frac{105.99 \text{ g Na}_2\text{CO}_3}{\text{mol}} = 53 \text{ g Na}_2\text{CO}_3$$

Dissolve 53 g Na_2CO_3 in water, dilute to 1.00 L.

$$17. \quad a. \quad 2.00 \text{ L} \times \frac{0.250 \text{ mol NaOH}}{\text{L}} \times \frac{40.00 \text{ g NaOH}}{\text{mol}} = 20.0 \text{ g NaOH}$$

Place 20.0 g NaOH in a 2-L volumetric flask; add water to dissolve the NaOH, and fill to the mark with water, mixing several times along the way.

$$b. \quad 2.00 \text{ L} \times \frac{0.250 \text{ mol NaOH}}{\text{L}} \times \frac{1 \text{ L stock}}{1.00 \text{ mol NaOH}} = 0.500 \text{ L}$$

Add 500. mL of 1.00 M NaOH stock solution to a 2-L volumetric flask; fill to the mark with water, mixing several times along the way.

$$c. \quad 2.00 \text{ L} \times \frac{0.100 \text{ mol K}_2\text{CrO}_4}{\text{L}} \times \frac{194.20 \text{ g K}_2\text{CrO}_4}{\text{mol K}_2\text{CrO}_4} = 38.8 \text{ g K}_2\text{CrO}_4$$

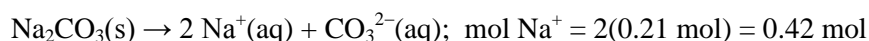
Similar to the solution made in part a, instead using 38.8 g K_2CrO_4 .

$$d. \quad 2.00 \text{ L} \times \frac{0.100 \text{ mol K}_2\text{CrO}_4}{\text{L}} \times \frac{1 \text{ L stock}}{1.75 \text{ mol K}_2\text{CrO}_4} = 0.114 \text{ L}$$

Similar to the solution made in part b, instead using 114 mL of the 1.75 M K_2CrO_4 stock solution.

$$18. \quad 75.0 \text{ mL} \times \frac{0.79 \text{ g}}{\text{mL}} \times \frac{1 \text{ mol}}{46.1 \text{ g}} = 1.3 \text{ mol C}_2\text{H}_5\text{OH}; \quad \text{molarity} = \frac{1.3 \text{ mol}}{0.250 \text{ L}} = 5.2 \text{ M C}_2\text{H}_5\text{OH}$$

$$19. \quad \text{Mol Na}_2\text{CO}_3 = 0.0700 \text{ L} \times \frac{3.0 \text{ mol Na}_2\text{CO}_3}{\text{L}} = 0.21 \text{ mol Na}_2\text{CO}_3$$



$$\text{Mol NaHCO}_3 = 0.0300 \text{ L} \times \frac{1.0 \text{ mol NaHCO}_3}{\text{L}} = 0.030 \text{ mol NaHCO}_3$$



$$M_{\text{Na}^+} = \frac{\text{total mol Na}^+}{\text{total volume}} = \frac{0.42 \text{ mol} + 0.030 \text{ mol}}{0.0700 \text{ L} + 0.0300 \text{ L}} = \frac{0.45 \text{ mol}}{0.1000 \text{ L}} = 4.5 \text{ M Na}^+$$

$$20. \quad 25.0 \text{ g (NH}_4)_2\text{SO}_4 \times \frac{1 \text{ mol}}{132.15 \text{ g}} = 1.89 \times 10^{-1} \text{ mol (NH}_4)_2\text{SO}_4$$

$$\text{Molarity} = \frac{1.89 \times 10^{-1} \text{ mol}}{100.0 \text{ mL}} \times \frac{1000 \text{ mL}}{\text{L}} = 1.89 \text{ M (NH}_4)_2\text{SO}_4$$

$$\begin{aligned} \text{Moles of (NH}_4)_2\text{SO}_4 \text{ in final solution} &= 10.00 \times 10^{-3} \text{ L} \times \frac{1.89 \text{ mol}}{\text{L}} \\ &= 1.89 \times 10^{-2} \text{ mol (NH}_4)_2\text{SO}_4 \end{aligned}$$

$$\text{Molarity of final solution} = \frac{1.89 \times 10^{-2} \text{ mol}}{(10.00 + 50.00) \text{ mL}} \times \frac{1000 \text{ mL}}{\text{L}} = 0.315 \text{ M (NH}_4)_2\text{SO}_4$$



$$21. \quad \text{Molar mass of NaOH} = 22.99 + 16.00 + 1.008 = 40.00 \text{ g/mol}$$

$$\text{Mass NaOH} = 0.2500 \text{ L} \times \frac{0.400 \text{ mol NaOH}}{\text{L}} \times \frac{40.00 \text{ g NaOH}}{\text{mol NaOH}} = 4.00 \text{ g NaOH}$$

22. Stock solution:

$$\begin{aligned} 1.584 \text{ g Mn}^{2+} \times \frac{1 \text{ mol Mn}^{2+}}{54.94 \text{ g Mn}^{2+}} &= 2.883 \times 10^{-2} \text{ mol Mn}^{2+}; \quad M = \frac{2.883 \times 10^{-2} \text{ mol}}{1.000 \text{ L}} \\ &= 2.883 \times 10^{-2} \text{ M Mn}^{2+} \end{aligned}$$

Solution A contains:

$$50.00 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{2.883 \times 10^{-2} \text{ mol}}{\text{L}} = 1.442 \times 10^{-3} \text{ mol Mn}^{2+}$$

$$\text{Molarity} = \frac{1.442 \times 10^{-3} \text{ mol}}{1000.0 \text{ mL}} \times \frac{1000 \text{ mL}}{\text{L}} = 1.442 \times 10^{-3} \text{ M Mn}^{2+}$$

Solution B contains:

$$10.00 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{1.442 \times 10^{-3} \text{ mol}}{\text{L}} = 1.442 \times 10^{-5} \text{ mol Mn}^{2+}$$

$$\text{Molarity} = \frac{1.442 \times 10^{-5} \text{ mol}}{0.2500 \text{ L}} = 5.768 \times 10^{-5} \text{ M Mn}^{2+}$$

Solution C contains:

$$10.00 \times 10^{-3} \text{ L} \times \frac{5.768 \times 10^{-5} \text{ mol}}{\text{L}} = 5.768 \times 10^{-7} \text{ mol Mn}^{2+}$$

$$\text{Molarity} = \frac{5.768 \times 10^{-7} \text{ mol}}{0.5000 \text{ L}} = 1.154 \times 10^{-6} \text{ M Mn}^{2+}$$

23. Stock solution = $\frac{10.0 \text{ mg}}{500.0 \text{ mL}} = \frac{10.0 \times 10^{-3} \text{ g}}{500.0 \text{ mL}} = \frac{2.00 \times 10^{-5} \text{ g steroid}}{\text{mL}}$

$$100.0 \times 10^{-6} \text{ L stock} \times \frac{1000 \text{ mL}}{\text{L}} \times \frac{2.00 \times 10^{-5} \text{ g steroid}}{\text{mL}} = 2.00 \times 10^{-6} \text{ g steroid}$$

This is diluted to a final volume of 100.0 mL.

$$\frac{2.00 \times 10^{-6} \text{ g steroid}}{100.0 \text{ mL}} \times \frac{1000 \text{ mL}}{\text{L}} \times \frac{1 \text{ mol steroid}}{336.4 \text{ g steroid}} = 5.95 \times 10^{-8} \text{ M steroid}$$

24. a. $M_{\text{Ca}(\text{NO}_3)_2} = \frac{0.100 \text{ mol Ca}(\text{NO}_3)_2}{0.100 \text{ L}} = 1.00 \text{ M}$

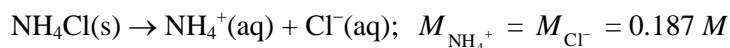


b. $M_{\text{Na}_2\text{SO}_4} = \frac{2.5 \text{ mol Na}_2\text{SO}_4}{1.25 \text{ L}} = 2.0 \text{ M}$



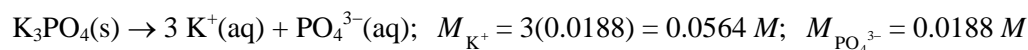
c. $5.00 \text{ g NH}_4\text{Cl} \times \frac{1 \text{ mol NH}_4\text{Cl}}{53.49 \text{ g NH}_4\text{Cl}} = 0.0935 \text{ mol NH}_4\text{Cl}$

$$M_{\text{NH}_4\text{Cl}} = \frac{0.0935 \text{ mol NH}_4\text{Cl}}{0.5000 \text{ L}} = 0.187 \text{ M}$$



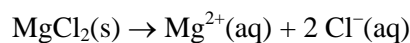
$$d. \quad 1.00 \text{ g K}_3\text{PO}_4 \times \frac{1 \text{ mol K}_3\text{PO}_4}{212.27 \text{ g}} = 4.71 \times 10^{-3} \text{ mol K}_3\text{PO}_4$$

$$M_{\text{K}_3\text{PO}_4} = \frac{4.71 \times 10^{-3} \text{ mol}}{0.2500 \text{ L}} = 0.0188 \text{ M}$$

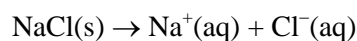


$$25. \quad \text{Mol solute} = \text{volume (L)} \times \text{molarity} \left(\frac{\text{mol}}{\text{L}} \right); \quad \text{AlCl}_3(\text{s}) \rightarrow \text{Al}^{3+}(\text{aq}) + 3 \text{ Cl}^-(\text{aq})$$

$$\text{Mol Cl}^- = 0.1000 \text{ L} \times \frac{0.30 \text{ mol AlCl}_3}{\text{L}} \times \frac{3 \text{ mol Cl}^-}{\text{mol AlCl}_3} = 9.0 \times 10^{-2} \text{ mol Cl}^-$$



$$\text{Mol Cl}^- = 0.0500 \text{ L} \times \frac{0.60 \text{ mol MgCl}_2}{\text{L}} \times \frac{2 \text{ mol Cl}^-}{\text{mol MgCl}_2} = 6.0 \times 10^{-2} \text{ mol Cl}^-$$



$$\text{Mol Cl}^- = 0.2000 \text{ L} \times \frac{0.40 \text{ mol NaCl}}{\text{L}} \times \frac{1 \text{ mol Cl}^-}{\text{mol NaCl}} = 8.0 \times 10^{-2} \text{ mol Cl}^-$$

100.0 mL of 0.30 M AlCl₃ contains the most moles of Cl⁻ ions.

$$26. \quad 10. \text{ g AgNO}_3 \times \frac{1 \text{ mol AgNO}_3}{169.9 \text{ g}} \times \frac{1 \text{ L}}{0.25 \text{ mol AgNO}_3} = 0.24 \text{ L} = 240 \text{ mL}$$

$$27. \quad a. \quad 5.0 \text{ ppb Hg in water} = \frac{5.0 \text{ ng Hg}}{\text{mL H}_2\text{O}} = \frac{5.0 \times 10^{-9} \text{ g Hg}}{\text{mL H}_2\text{O}}$$

$$\frac{5.0 \times 10^{-9} \text{ g Hg}}{\text{mL}} \times \frac{1 \text{ mol Hg}}{200.6 \text{ g Hg}} \times \frac{1000 \text{ mL}}{\text{L}} = 2.5 \times 10^{-8} \text{ M Hg}$$

$$b. \quad \frac{1.0 \times 10^{-9} \text{ g CHCl}_3}{\text{mL}} \times \frac{1 \text{ mol CHCl}_3}{119.4 \text{ g CHCl}_3} \times \frac{1000 \text{ mL}}{\text{L}} = 8.4 \times 10^{-9} \text{ M CHCl}_3$$

$$c. \quad 10.0 \text{ ppm As} = \frac{10.0 \text{ } \mu\text{g As}}{\text{mL}} = \frac{10.0 \times 10^{-6} \text{ g As}}{\text{mL}}$$

$$\frac{10.0 \times 10^{-6} \text{ g As}}{\text{mL}} \times \frac{1 \text{ mol As}}{74.92 \text{ g As}} \times \frac{1000 \text{ mL}}{\text{L}} = 1.33 \times 10^{-4} \text{ M As}$$

$$d. \frac{0.10 \times 10^{-6} \text{ g DDT}}{\text{mL}} \times \frac{1 \text{ mol DDT}}{354.5 \text{ g DDT}} \times \frac{1000 \text{ mL}}{\text{L}} = 2.8 \times 10^{-7} \text{ M DDT}$$

28. We want 100.0 mL of each standard. To make the 100. ppm standard:

$$\frac{100. \mu\text{g Cu}}{\text{mL}} \times 100.0 \text{ mL solution} = 1.00 \times 10^4 \mu\text{g Cu needed}$$

$$1.00 \times 10^4 \mu\text{g Cu} \times \frac{1 \text{ mL stock}}{1000.0 \mu\text{g Cu}} = 10.0 \text{ mL of stock solution}$$

Therefore, to make 100.0 mL of 100. ppm solution, transfer 10.0 mL of the 1000.0 ppm stock solution to a 100-mL volumetric flask, and dilute to the mark.

Similarly:

75.0 ppm standard, dilute 7.50 mL of the 1000.0 ppm stock to 100.0 mL.

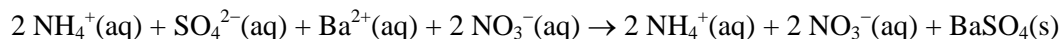
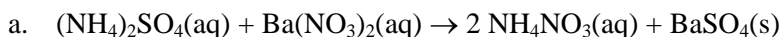
50.0 ppm standard, dilute 5.00 mL of the 1000.0 ppm stock to 100.0 mL.

25.0 ppm standard, dilute 2.50 mL of the 1000.0 ppm stock to 100.0 mL.

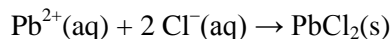
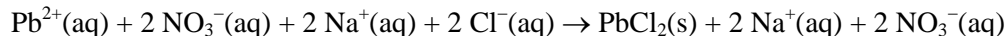
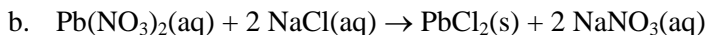
10.0 ppm standard, dilute 1.00 mL of the 1000.0 ppm stock to 100.0 mL.

Precipitation Reactions

29. For the following answers, the balanced molecular equation is first, followed by the complete ionic equation, and then the net ionic equation.

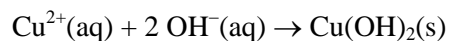
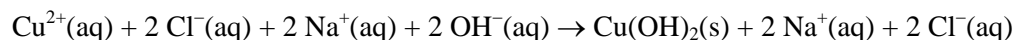
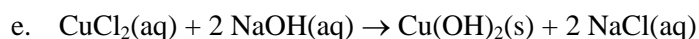


$\text{Ba}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{BaSO}_4(\text{s})$ is the net ionic equation (spectator ions omitted).

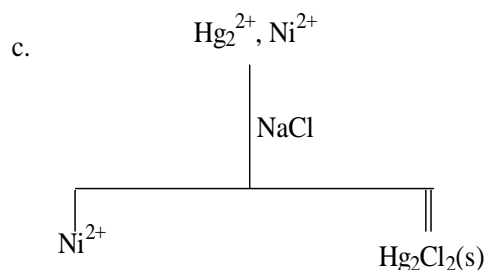
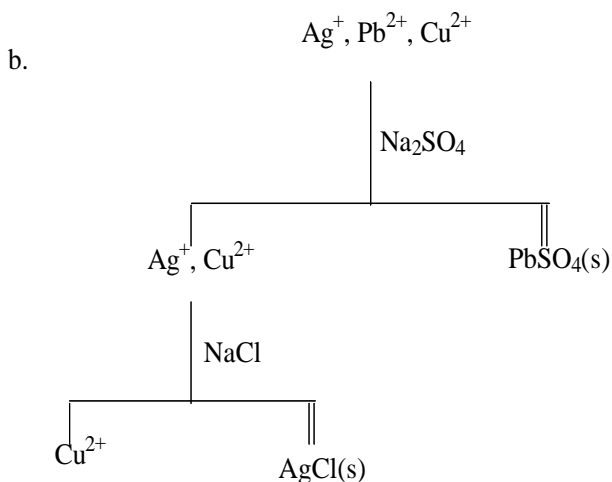
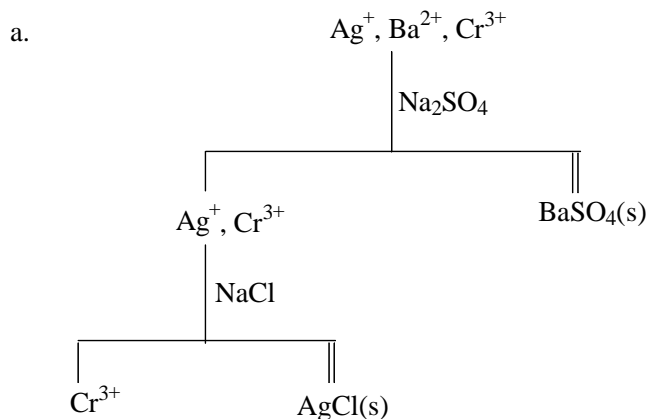


c. The possible products, potassium phosphate and sodium nitrate, are both soluble in water. Therefore, no reaction occurs.

d. No reaction occurs because all possible products are soluble.



30. The following schemes show reagents to add in order to precipitate one ion at a time. In each scheme, NaOH can be added to precipitate the last remaining ion.



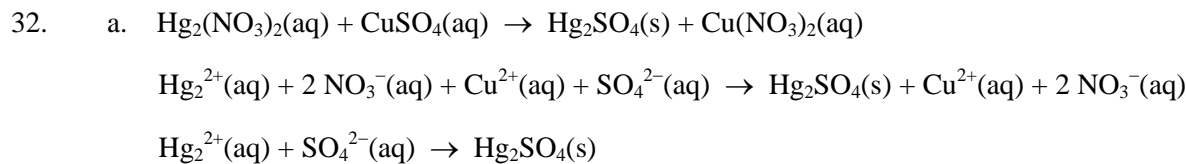
31. Use Table 4.1 to predict the solubility of the possible products.

a. Possible products = Hg_2SO_4 and $\text{Cu}(\text{NO}_3)_2$; precipitate = Hg_2SO_4

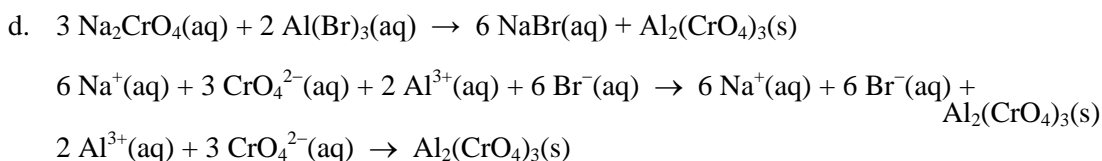
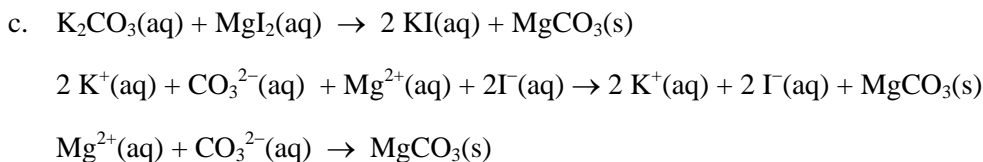
b. Possible products = NiCl_2 and $\text{Ca}(\text{NO}_3)_2$; both salts are soluble so no precipitate forms.

c. Possible products = KI and MgCO_3 ; precipitate = MgCO_3

d. Possible products = NaBr and $\text{Al}_2(\text{CrO}_4)_3$; precipitate = $\text{Al}_2(\text{CrO}_4)_3$



b. No reaction occurs since both possible products are soluble.



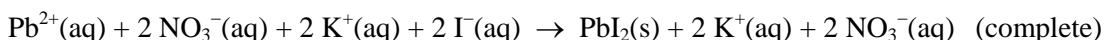
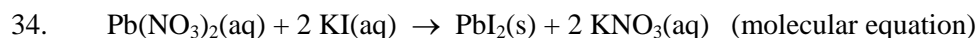
33. Use the solubility rules in Table 4.1. Some soluble bromides by Rule 2 would be NaBr, KBr, and NH_4Br (there are others). The insoluble bromides by Rule 3 would be AgBr, PbBr_2 , and Hg_2Br_2 . Similar reasoning is used for the other parts to this problem.

Sulfates: Na_2SO_4 , K_2SO_4 , and $(\text{NH}_4)_2\text{SO}_4$ (and others) would be soluble, and BaSO_4 , CaSO_4 , and PbSO_4 (or Hg_2SO_4) would be insoluble.

Hydroxides: NaOH, KOH, $\text{Ca}(\text{OH})_2$ (and others) would be soluble, and $\text{Al}(\text{OH})_3$, $\text{Fe}(\text{OH})_3$, and $\text{Cu}(\text{OH})_2$ (and others) would be insoluble.

Phosphates: Na_3PO_4 , K_3PO_4 , $(\text{NH}_4)_3\text{PO}_4$ (and others) would be soluble, and Ag_3PO_4 , $\text{Ca}_3(\text{PO}_4)_2$, and FePO_4 (and others) would be insoluble.

Lead: PbCl_2 , PbBr_2 , PbI_2 , $\text{Pb}(\text{OH})_2$, PbSO_4 , and PbS (and others) would be insoluble. $\text{Pb}(\text{NO}_3)_2$ would be a soluble Pb^{2+} salt.

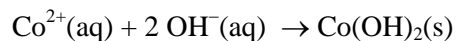
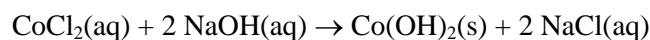


The 1.0 mol of Pb^{2+} ions would react with the 2.0 mol of I^- ions to form 1.0 mol of the PbI_2 precipitate. Even though the Pb^{2+} and I^- ions are removed, the spectator ions K^+ and NO_3^- are still present. The solution above the precipitate will conduct electricity because there are plenty of charge carriers (K^+ and NO_3^- ions) present in solution.

35. a. When $\text{CuSO}_4(\text{aq})$ is added to $\text{Na}_2\text{S}(\text{aq})$, the precipitate that forms is $\text{CuS}(\text{s})$. Therefore, Na^+ (the gray spheres) and SO_4^{2-} (the bluish green spheres) are the spectator ions.



- b. When $\text{CoCl}_2(\text{aq})$ is added to $\text{NaOH}(\text{aq})$, the precipitate that forms is $\text{Co}(\text{OH})_2(\text{s})$. Therefore, Na^+ (the gray spheres) and Cl^- (the green spheres) are the spectator ions.



- c. When $\text{AgNO}_3(\text{aq})$ is added to $\text{KI}(\text{aq})$, the precipitate that forms is $\text{AgI}(\text{s})$. Therefore, K^+ (the red spheres) and NO_3^- (the blue spheres) are the spectator ions.



36. $\text{XCl}_2(\text{aq}) + 2 \text{AgNO}_3(\text{aq}) \rightarrow 2 \text{AgCl}(\text{s}) + \text{X}(\text{NO}_3)_2(\text{aq})$

$$1.38 \text{ g AgCl} \times \frac{1 \text{ mol AgCl}}{143.4 \text{ g}} \times \frac{1 \text{ mol XCl}_2}{2 \text{ mol AgCl}} = 4.81 \times 10^{-3} \text{ mol XCl}_2$$

$$\frac{1.00 \text{ g XCl}_2}{4.91 \times 10^{-3} \text{ mol XCl}_2} = 208 \text{ g/mol}; \quad x + 2(35.45) = 208, \quad x = 137 \text{ g/mol}$$

The metal X is barium (Ba).

37. $\text{M}_2\text{SO}_4(\text{aq}) + \text{CaCl}_2(\text{aq}) \rightarrow \text{CaSO}_4(\text{s}) + 2 \text{MCl}(\text{aq})$

$$1.36 \text{ g CaSO}_4 \times \frac{1 \text{ mol CaSO}_4}{136.15 \text{ g CaSO}_4} \times \frac{1 \text{ mol M}_2\text{SO}_4}{\text{mol CaSO}_4} = 9.99 \times 10^{-3} \text{ mol M}_2\text{SO}_4$$

From the problem, 1.42 g M_2SO_4 was reacted, so:

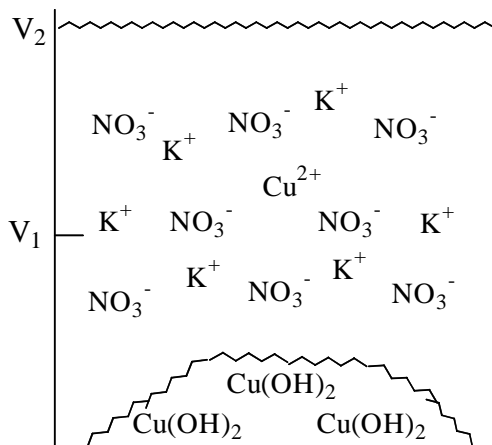
$$\text{molar mass} = \frac{1.42 \text{ g M}_2\text{SO}_4}{9.99 \times 10^{-3} \text{ mol M}_2\text{SO}_4} = 142 \text{ g/mol}$$

$$142 \text{ amu} = 2(\text{atomic mass M}) + 32.07 + 4(16.00), \quad \text{atomic mass M} = 23 \text{ amu}$$

From periodic table, M is Na (sodium).

38. a. $\text{Cu}(\text{NO}_3)_2(\text{aq}) + 2 \text{KOH}(\text{aq}) \rightarrow \text{Cu}(\text{OH})_2(\text{s}) + 2 \text{KNO}_3(\text{aq})$

Solution A contains $2.00 \text{ L} \times 2.00 \text{ mol/L} = 4.00 \text{ mol Cu}(\text{NO}_3)_2$, and solution B contains $2.00 \text{ L} \times 3.00 \text{ mol/L} = 6.00 \text{ mol KOH}$. Let's assume in our picture that we have 4 formula units of $\text{Cu}(\text{NO}_3)_2$ (4 Cu^{2+} ions and 8 NO_3^- ions) and 6 formula units of KOH (6 K^+ ions and 6 OH^- ions). With 4 Cu^{2+} ions and 6 OH^- ions present, then OH^- is limiting. One Cu^{2+} ion remains as 3 $\text{Cu}(\text{OH})_2(\text{s})$ formula units form as precipitate. The following illustration summarizes the ions that remain in solution and the relative amount of precipitate that forms. Note that K^+ and NO_3^- ions are spectator ions. In the illustration, V_1 is the volume of solution A or B and V_2 is the volume of the combined solutions with $V_2 = 2V_1$. The illustration exaggerates the amount of precipitate that would actually form.



- b. The spectator ion concentrations will be one-half the original spectator ion concentrations in the individual beakers because the volume was doubled. Or using moles, $M_{\text{K}^+} =$

$$\frac{6.00 \text{ mol K}^+}{4.00 \text{ L}} = 1.50 \text{ M} \text{ and } M_{\text{NO}_3^-} = \frac{8.00 \text{ mol NO}_3^-}{4.00 \text{ L}} = 2.00 \text{ M. The concentration of}$$

OH^- ions will be zero because OH^- is the limiting reagent. From the drawing, the number of Cu^{2+} ions will decrease by a factor of four as the precipitate forms. Because the volume of solution doubled, the concentration of Cu^{2+} ions will decrease by a factor of eight after the two beakers are mixed:

$$M_{\text{Cu}^{2+}} = 2.00 \text{ M} \left(\frac{1}{8} \right) = 0.250 \text{ M}$$

Alternately, one could certainly use moles to solve for $M_{\text{Cu}^{2+}}$:

$$\text{Mol Cu}^{2+} \text{ reacted} = 2.00 \text{ L} \times \frac{3.00 \text{ mol OH}^-}{\text{L}} \times \frac{1 \text{ mol Cu}^{2+}}{2 \text{ mol OH}^-} = 3.00 \text{ mol Cu}^{2+} \text{ reacted}$$

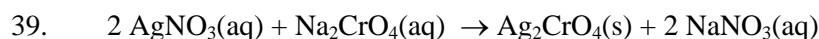
$$\text{Mol Cu}^{2+} \text{ present initially} = 2.00 \text{ L} \times \frac{2.00 \text{ mol Cu}^{2+}}{\text{L}} = 4.00 \text{ mol Cu}^{2+} \text{ present initially}$$

$$\text{Excess Cu}^{2+} \text{ present after reaction} = 4.00 \text{ mol} - 3.00 \text{ mol} = 1.00 \text{ mol Cu}^{2+} \text{ excess}$$

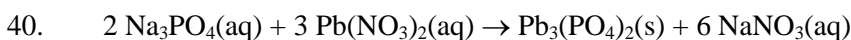
$$M_{\text{Cu}^{2+}} = \frac{1.00 \text{ mol Cu}^{2+}}{2.00 \text{ L} + 2.00 \text{ L}} = 0.250 \text{ M}$$

$$\text{Mass of precipitate} = 6.00 \text{ mol KOH} \times \frac{1 \text{ mol Cu(OH)}_2}{2 \text{ mol KOH}} \times \frac{97.57 \text{ g Cu(OH)}_2}{\text{mol Cu(OH)}_2}$$

$$\text{Mass of precipitate} = 293 \text{ g Cu(OH)}_2$$

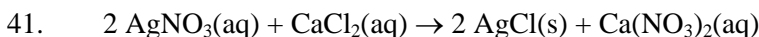


$$0.0750 \text{ L} \times \frac{0.100 \text{ mol AgNO}_3}{\text{L}} \times \frac{1 \text{ mol Na}_2\text{CrO}_4}{2 \text{ mol AgNO}_3} \times \frac{161.98 \text{ g Na}_2\text{CrO}_4}{\text{mol Na}_2\text{CrO}_4} = 0.607 \text{ g Na}_2\text{CrO}_4$$



$$0.1500 \text{ L} \times \frac{0.250 \text{ mol Pb}(\text{NO}_3)_2}{\text{L}} \times \frac{2 \text{ mol Na}_3\text{PO}_4}{3 \text{ mol Pb}(\text{NO}_3)_2} \times \frac{1 \text{ L Na}_3\text{PO}_4}{0.100 \text{ mol Na}_3\text{PO}_4} = 0.250 \text{ L}$$

$$= 250. \text{ mL Na}_3\text{PO}_4$$



$$\text{Mol AgNO}_3 = 0.1000 \text{ L} \times \frac{0.20 \text{ mol AgNO}_3}{\text{L}} = 0.020 \text{ mol AgNO}_3$$

$$\text{Mol CaCl}_2 = 0.1000 \text{ L} \times \frac{0.15 \text{ mol CaCl}_2}{\text{L}} = 0.015 \text{ mol CaCl}_2$$

The required mol AgNO₃ to mol CaCl₂ ratio is 2 : 1 (from the balanced equation). The actual mole ratio present is 0.020/0.015 = 1.3 (1.3 : 1). Therefore, AgNO₃ is the limiting reagent.

$$\text{Mass AgCl} = 0.020 \text{ mol AgNO}_3 \times \frac{1 \text{ mol AgCl}}{1 \text{ mol AgNO}_3} \times \frac{143.4 \text{ g AgCl}}{\text{mol AgCl}} = 2.9 \text{ g AgCl}$$

The net ionic equation is $\text{Ag}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \rightarrow \text{AgCl}(\text{s})$. The ions remaining in solution are the unreacted Cl⁻ ions and the spectator ions, NO₃⁻ and Ca²⁺ (all Ag⁺ is used up in forming AgCl). The moles of each ion present initially (before reaction) can be easily determined from the moles of each reactant. 0.020 mol AgNO₃ dissolves to form 0.020 mol Ag⁺ and 0.020 mol NO₃⁻. 0.015 mol CaCl₂ dissolves to form 0.015 mol Ca²⁺ and 2(0.015) = 0.030 mol Cl⁻.

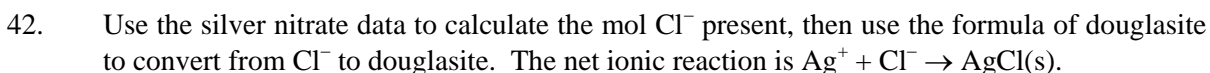
$$\text{Mol unreacted Cl}^- = 0.030 \text{ mol Cl}^- \text{ initially} - 0.020 \text{ mol Cl}^- \text{ reacted}$$

$$\text{Mol unreacted Cl}^- = 0.010 \text{ mol Cl}^-$$

$$M_{\text{Cl}^-} = \frac{0.010 \text{ mol Cl}^-}{\text{total volume}} = \frac{0.010 \text{ mol Cl}^-}{0.1000 \text{ L} + 0.1000 \text{ L}} = 0.050 \text{ M Cl}^-$$

The molarity of the spectator ions are:

$$M_{\text{NO}_3^-} = \frac{0.020 \text{ mol NO}_3^-}{0.2000 \text{ L}} = 0.10 \text{ M NO}_3^-; \quad M_{\text{Ca}^{2+}} = \frac{0.015 \text{ mol Ca}^{2+}}{0.2000 \text{ L}} = 0.075 \text{ M Ca}^{2+}$$



$$0.03720 \text{ L} \times \frac{0.1000 \text{ mol Ag}^+}{\text{L}} \times \frac{1 \text{ mol Cl}^-}{\text{mol Ag}^+} \times \frac{1 \text{ mol douglasite}}{4 \text{ mol Cl}^-} \times \frac{311.88 \text{ g douglasite}}{\text{mol}}$$

$$= 0.2900 \text{ g douglasite}$$

$$\text{Mass \% douglasite} = \frac{0.2900\text{g}}{0.4550\text{g}} \times 100 = 63.74\%$$

43. All the sulfur in BaSO_4 came from the saccharin. The conversion from BaSO_4 to saccharin uses the molar masses and formulas of each compound.

$$0.5032\text{ g BaSO}_4 \times \frac{32.07\text{g S}}{233.4\text{g BaSO}_4} \times \frac{183.9\text{g saccharin}}{32.07\text{g S}} = 0.3949\text{ g saccharin}$$

$$\frac{\text{Average mass}}{\text{Tablet}} = \frac{0.3949\text{g}}{10\text{ tablets}} = \frac{3.949 \times 10^{-2}\text{ g}}{\text{tablet}} = \frac{39.49\text{ mg}}{\text{tablet}}$$

$$\text{Average mass \%} = \frac{0.3949\text{g saccharin}}{0.5894\text{g}} \times 100 = 67.00\% \text{ saccharin by mass}$$

44. All the Tl in TII came from Tl in Tl_2SO_4 . The conversion from TII to Tl_2SO_4 uses the molar masses and formulas of each compound.

$$0.1824\text{ g TII} \times \frac{204.4\text{ g Tl}}{331.3\text{ g TII}} \times \frac{504.9\text{ g Tl}_2\text{SO}_4}{408.8\text{ g Tl}} = 0.1390\text{ g Tl}_2\text{SO}_4$$

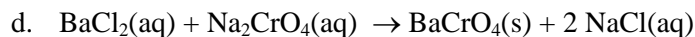
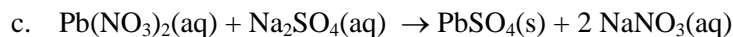
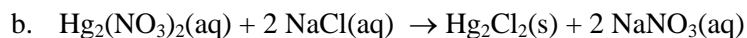
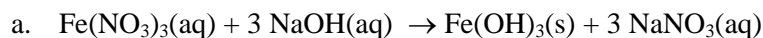
$$\text{Mass \% Tl}_2\text{SO}_4 = \frac{0.1390\text{g Tl}_2\text{SO}_4}{9.486\text{g pesticide}} \times 100 = 1.465\% \text{ Tl}_2\text{SO}_4$$

45. Use aluminum in the formulas to convert from mass of $\text{Al}(\text{OH})_3$ to mass of $\text{Al}_2(\text{SO}_4)_3$ in the mixture.

$$0.107\text{ g Al}(\text{OH})_3 \times \frac{1\text{ mol Al}(\text{OH})_3}{78.00\text{ g}} \times \frac{1\text{ mol Al}^{3+}}{\text{mol Al}(\text{OH})_3} \times \frac{1\text{ mol Al}_2(\text{SO}_4)_3}{2\text{ mol Al}^{3+}} \times \frac{342.17\text{g Al}_2(\text{SO}_4)_3}{\text{mol Al}_2(\text{SO}_4)_3} = 0.235\text{ g Al}_2(\text{SO}_4)_3$$

$$\text{Mass \% Al}_2(\text{SO}_4)_3 = \frac{0.235\text{ g}}{1.45\text{ g}} \times 100 = 16.2\%$$

46. There are many acceptable choices for spectator ions. We will generally choose Na^+ and NO_3^- as the spectator ions because sodium salts and nitrate salts are usually soluble in water.



47. Because a precipitate formed with Na_2SO_4 , the possible cations are Ba^{2+} , Pb^{2+} , Hg_2^{2+} , and Ca^{2+} (from the solubility rules). Because no precipitate formed with KCl , Pb^{2+} , and Hg_2^{2+} cannot be present. Because both Ba^{2+} and Ca^{2+} form soluble chlorides and soluble hydroxides, both these cations could be present. Therefore, the cations could be Ba^{2+} and Ca^{2+} (by the solubility rules in Table 4.1). For students who do a more rigorous study of solubility, Sr^{2+} could also be a possible cation (it forms an insoluble sulfate salt, whereas the chloride and hydroxide salts of strontium are soluble).

Acid-Base Reactions

48. Strong bases contain the hydroxide ion, OH^- . The reaction that occurs is $\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$.

$$0.0120 \text{ L} \times \frac{0.150 \text{ mol H}^+}{\text{L}} \times \frac{1 \text{ mol OH}^-}{\text{mol H}^+} = 1.80 \times 10^{-3} \text{ mol OH}^-$$

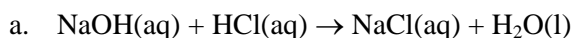
The 30.0 mL of the unknown strong base contains $1.80 \times 10^{-3} \text{ mol OH}^-$.

$$\frac{1.8 \times 10^{-3} \text{ mol OH}^-}{0.0300 \text{ L}} = 0.0600 \text{ M OH}^-$$

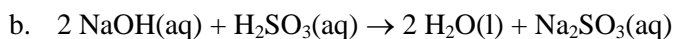
The unknown base concentration is one-half the concentration of OH^- ions produced from the base, so the base must contain 2 OH^- in each formula unit. The three soluble strong bases that have 2 OH^- ions in the formula are $\text{Ca}(\text{OH})_2$, $\text{Sr}(\text{OH})_2$, and $\text{Ba}(\text{OH})_2$. These are all possible identities for the strong base.

49. If we begin with 50.00 mL of 0.100 M NaOH, then:

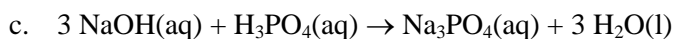
$$50.00 \times 10^{-3} \text{ L} \times \frac{0.100 \text{ mol}}{\text{L}} = 5.00 \times 10^{-3} \text{ mol NaOH to be neutralized.}$$



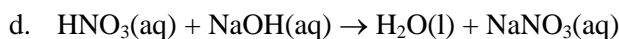
$$5.00 \times 10^{-3} \text{ mol NaOH} \times \frac{1 \text{ mol HCl}}{\text{mol NaOH}} \times \frac{1 \text{ L soln}}{0.100 \text{ mol}} = 5.00 \times 10^{-2} \text{ L or 50.0 mL}$$



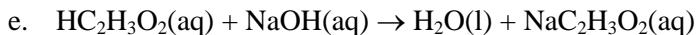
$$5.00 \times 10^{-3} \text{ mol NaOH} \times \frac{1 \text{ mol H}_2\text{SO}_3}{2 \text{ mol NaOH}} \times \frac{1 \text{ L soln}}{0.100 \text{ mol H}_2\text{SO}_3} = 2.50 \times 10^{-2} \text{ L or 25.0 mL}$$



$$5.00 \times 10^{-3} \text{ mol NaOH} \times \frac{1 \text{ mol H}_3\text{PO}_4}{3 \text{ mol NaOH}} \times \frac{1 \text{ L soln}}{0.200 \text{ mol H}_3\text{PO}_4} = 8.33 \times 10^{-3} \text{ L or 8.33 mL}$$

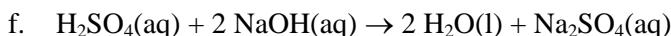


$$5.00 \times 10^{-3} \text{ mol NaOH} \times \frac{1 \text{ mol HNO}_3}{\text{mol NaOH}} \times \frac{1 \text{ L soln}}{0.150 \text{ mol HNO}_3} = 3.33 \times 10^{-2} \text{ L or } 33.3 \text{ mL}$$

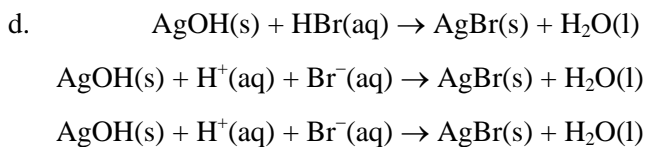
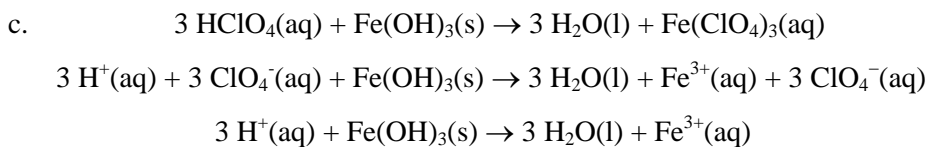
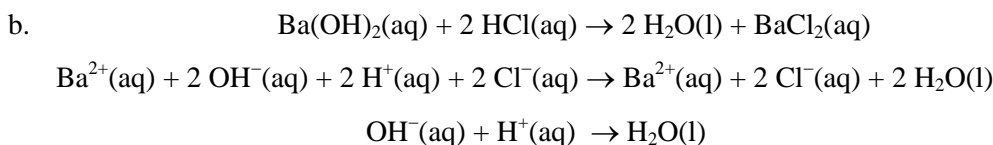
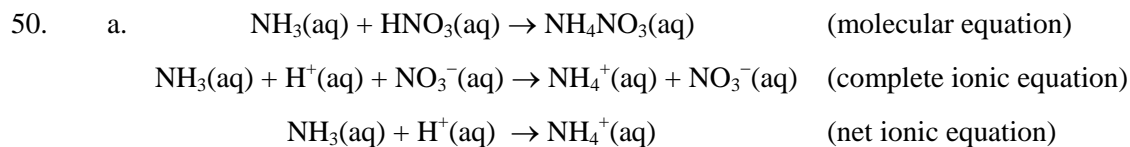


$$5.00 \times 10^{-3} \text{ mol NaOH} \times \frac{1 \text{ mol HC}_2\text{H}_3\text{O}_2}{\text{mol NaOH}} \times \frac{1 \text{ L soln}}{0.200 \text{ mol HC}_2\text{H}_3\text{O}_2} = 2.50 \times 10^{-2} \text{ L}$$

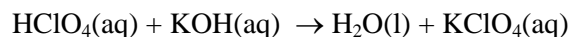
or 25.0 mL



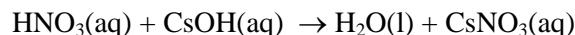
$$5.00 \times 10^{-3} \text{ mol NaOH} \times \frac{1 \text{ mol H}_2\text{SO}_4}{2 \text{ mol NaOH}} \times \frac{1 \text{ L soln}}{0.300 \text{ mol H}_2\text{SO}_4} = 8.33 \times 10^{-3} \text{ L or } 8.33 \text{ mL}$$



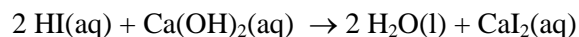
51. a. Perchloric acid reacted with potassium hydroxide is a possibility.



b. Nitric acid reacted with cesium hydroxide is a possibility.



c. Hydroiodic acid reacted with calcium hydroxide is a possibility.



52. We get the empirical formula from the elemental analysis. Out of 100.00 g carminic acid there are:

$$53.66 \text{ g C} \times \frac{1 \text{ mol C}}{12.011 \text{ g C}} = 4.468 \text{ mol C}; \quad 4.09 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 4.06 \text{ mol H}$$

$$42.25 \text{ g O} \times \frac{1 \text{ mol O}}{15.999 \text{ g O}} = 2.641 \text{ mol O}$$

Dividing the moles by the smallest number gives:

$$\frac{4.468}{2.641} = 1.692; \quad \frac{4.06}{2.641} = 1.54$$

These numbers don't give obvious mole ratios. Let's determine the mol C to mol H ratio:

$$\frac{4.468}{4.06} = 1.10 = \frac{11}{10}$$

So let's try $\frac{4.06}{10} = 0.406$ as a common factor: $\frac{4.468}{0.406} = 11.0$; $\frac{4.06}{0.406} = 10.0$; $\frac{2.641}{0.406} = 6.50$

Therefore, $\text{C}_{22}\text{H}_{20}\text{O}_{13}$ is the empirical formula.

We can get molar mass from the titration data. The balanced reaction is $\text{HA}(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l}) + \text{A}^-(\text{aq})$ where HA is an abbreviation for carminic acid, an acid with one acidic proton (H^+).

$$18.02 \times 10^{-3} \text{ L soln} \times \frac{0.0406 \text{ mol NaOH}}{\text{L soln}} \times \frac{1 \text{ mol carminic acid}}{\text{mol NaOH}} = 7.32 \times 10^{-4} \text{ mol carminic acid}$$

$$\text{Molar mass} = \frac{0.3602 \text{ g}}{7.32 \times 10^{-4} \text{ mol}} = \frac{492 \text{ g}}{\text{mol}}$$

The empirical formula mass of $\text{C}_{22}\text{H}_{20}\text{O}_{13} \approx 22(12) + 20(1) + 13(16) = 492 \text{ g}$.

Therefore, the molecular formula of carminic acid is also $\text{C}_{22}\text{H}_{20}\text{O}_{13}$.

53. HCl and HNO_3 are strong acids; $\text{Ca}(\text{OH})_2$ and RbOH are strong bases. The net ionic equation that occurs is $\text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l})$.

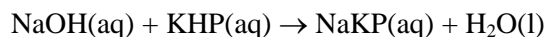
$$\begin{aligned} \text{Mol H}^+ &= 0.0500 \text{ L} \times \frac{0.100 \text{ mol HCl}}{\text{L}} \times \frac{1 \text{ mol H}^+}{\text{mol HCl}} + \\ &0.1000 \text{ L} \times \frac{0.200 \text{ mol HNO}_3}{\text{L}} \times \frac{1 \text{ mol H}^+}{\text{mol HNO}_3} = 0.00500 + 0.0200 = 0.0250 \text{ mol H}^+ \end{aligned}$$

$$\begin{aligned} \text{Mol OH}^- &= 0.5000 \text{ L} \times \frac{0.0100 \text{ mol Ca(OH)}_2}{\text{L}} \times \frac{2 \text{ mol OH}^-}{\text{mol Ca(OH)}_2} + \\ &0.2000 \text{ L} \times \frac{0.100 \text{ mol RbOH}}{\text{L}} \times \frac{1 \text{ mol OH}^-}{\text{mol RbOH}} = 0.0100 + 0.0200 = 0.0300 \text{ mol OH}^- \end{aligned}$$

We have an excess of OH^- , so the solution is basic (not neutral). The moles of excess $\text{OH}^- = 0.0300 \text{ mol OH}^-$ initially $- 0.0250 \text{ mol OH}^-$ reacted (with H^+) $= 0.0050 \text{ mol OH}^-$ excess.

$$M_{\text{OH}^-} = \frac{0.0050 \text{ mol OH}^-}{(0.05000 + 0.1000 + 0.5000 + 0.2000) \text{ L}} = \frac{0.0050 \text{ mol}}{0.8500 \text{ L}} = 5.9 \times 10^{-3} \text{ M}$$

54. Because KHP is a monoprotic acid, the reaction is (KHP is an abbreviation for potassium hydrogen phthalate):



$$0.1082 \text{ g KHP} \times \frac{1 \text{ mol KHP}}{204.22 \text{ g KHP}} \times \frac{1 \text{ mol NaOH}}{\text{mol KHP}} = 5.298 \times 10^{-4} \text{ mol NaOH}$$

There is $5.298 \times 10^{-4} \text{ mol}$ of sodium hydroxide in 34.67 mL of solution. Therefore, the concentration of sodium hydroxide is:

$$\frac{5.298 \times 10^{-4} \text{ mol}}{34.67 \times 10^{-3} \text{ L}} = 1.528 \times 10^{-2} \text{ M NaOH}$$

55. $\text{HC}_2\text{H}_3\text{O}_2(\text{aq}) + \text{NaOH}(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l}) + \text{NaC}_2\text{H}_3\text{O}_2(\text{aq})$

$$\begin{aligned} \text{a. } 16.58 \times 10^{-3} \text{ L soln H} &\times \frac{0.5062 \text{ mol NaOH}}{\text{L soln}} \times \frac{1 \text{ mol acetic acid}}{\text{mol NaOH}} \\ &= 8.393 \times 10^{-3} \text{ mol acetic acid} \end{aligned}$$

$$\text{Concentration of acetic acid} = \frac{8.393 \times 10^{-3} \text{ mol}}{0.01000 \text{ L}} = 0.8393 \text{ M HC}_2\text{H}_3\text{O}_2$$

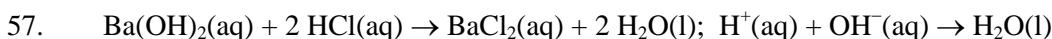
$$\text{b. If we have 1.000 L of solution: total mass} = 1000. \text{ mL} \times \frac{1.006 \text{ g}}{\text{mL}} = 1006 \text{ g solution}$$

$$\text{Mass of HC}_2\text{H}_3\text{O}_2 = 0.8393 \text{ mol} \times \frac{60.052 \text{ g}}{\text{mol}} = 50.40 \text{ g HC}_2\text{H}_3\text{O}_2$$

$$\text{Mass \% acetic acid} = \frac{50.40 \text{ g}}{1006 \text{ g}} \times 100 = 5.010\%$$

56. $39.47 \times 10^{-3} \text{ L HCl} \times \frac{0.0984 \text{ mol HCl}}{\text{L}} \times \frac{1 \text{ mol NH}_3}{\text{mol HCl}} = 3.88 \times 10^{-3} \text{ mol NH}_3$

$$\text{Molarity of NH}_3 = \frac{3.88 \times 10^{-3} \text{ mol}}{50.00 \times 10^{-3} \text{ L}} = 0.0776 \text{ M NH}_3$$

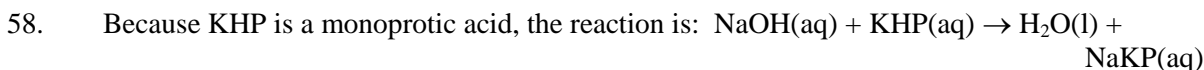


$$75.0 \times 10^{-3} \text{ L} \times \frac{0.250 \text{ mol HCl}}{\text{L}} = 1.88 \times 10^{-2} \text{ mol HCl} = 1.88 \times 10^{-2} \text{ mol H}^+ + 1.88 \times 10^{-2} \text{ mol Cl}^-$$

$$225.0 \times 10^{-3} \text{ L} \times \frac{0.0550 \text{ mol Ba(OH)}_2}{\text{L}} = 1.24 \times 10^{-2} \text{ mol Ba(OH)}_2 = 1.24 \times 10^{-2} \text{ mol Ba}^{2+} + 2.48 \times 10^{-2} \text{ mol OH}^-$$

The net ionic equation requires a 1 : 1 mol ratio between OH^- and H^+ . The actual mol OH^- to mol H^+ ratio is greater than 1 : 1, so OH^- is in excess. Because 1.88×10^{-2} mol OH^- will be neutralized by the H^+ , we have $(2.48 - 1.88) \times 10^{-2} = 0.60 \times 10^{-2}$ mol OH^- in excess.

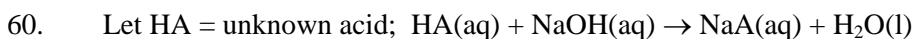
$$M_{\text{OH}^-} = \frac{\text{mol OH}^- \text{ excess}}{\text{total volume}} = \frac{6.0 \times 10^{-3} \text{ mol OH}^-}{0.0750 \text{ L} + 0.2250 \text{ L}} = 2.0 \times 10^{-2} \text{ M OH}^-$$



$$\begin{aligned} \text{Mass KHP} &= 0.02046 \text{ L NaOH} \times \frac{0.1000 \text{ mol NaOH}}{\text{L NaOH}} \times \frac{1 \text{ mol KHP}}{\text{mol NaOH}} \times \frac{204.22 \text{ g KHP}}{\text{mol KHP}} \\ &= 0.4178 \text{ g KHP} \end{aligned}$$

59. The acid is a diprotic acid (H_2A) meaning that it has two H^+ ions in the formula to donate to a base. The reaction is $\text{H}_2\text{A}(\text{aq}) + 2 \text{NaOH}(\text{aq}) \rightarrow 2 \text{H}_2\text{O}(\text{l}) + \text{Na}_2\text{A}(\text{aq})$, where A^{2-} is what is left over from the acid formula when the two protons (H^+ ions) are reacted.

For the HCl reaction, the base has the ability to accept two protons. The most common examples are Ca(OH)_2 , Sr(OH)_2 , and Ba(OH)_2 . A possible reaction would be $2 \text{HCl}(\text{aq}) + \text{Ca(OH)}_2(\text{aq}) \rightarrow 2 \text{H}_2\text{O}(\text{l}) + \text{CaCl}_2(\text{aq})$.



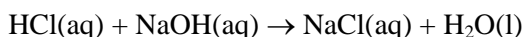
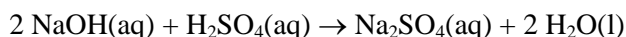
$$\text{Mol HA present} = 0.0250 \text{ L} \times \frac{0.500 \text{ mol NaOH}}{\text{L}} \times \frac{1 \text{ mol HA}}{1 \text{ mol NaOH}} = 0.0125 \text{ mol HA}$$

$$\frac{x \text{ g HA}}{\text{mol HA}} = \frac{2.20 \text{ g HA}}{0.0125 \text{ mol HA}}, \quad x = \text{molar mass of HA} = 176 \text{ g/mol}$$

Empirical formula mass $\approx 3(12) + 4(1) + 3(16) = 88 \text{ g/mol}$.

Because $176/88 = 2.0$, the molecular formula is $(\text{C}_3\text{H}_4\text{O}_3)_2 = \text{C}_6\text{H}_8\text{O}_6$.

61. The pertinent reactions are:



$$\text{Amount of NaOH added} = 0.0500 \text{ L} \times \frac{0.213 \text{ mol}}{\text{L}} = 1.07 \times 10^{-2} \text{ mol NaOH}$$

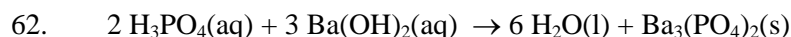
Amount of NaOH neutralized by HCl:

$$0.01321 \text{ L HCl} \times \frac{0.103 \text{ mol HCl}}{\text{L HCl}} \times \frac{1 \text{ mol NaOH}}{\text{mol HCl}} = 1.36 \times 10^{-3} \text{ mol NaOH}$$

The difference, 9.3×10^{-3} mol, is the amount of NaOH neutralized by the sulfuric acid.

$$9.3 \times 10^{-3} \text{ mol NaOH} \times \frac{1 \text{ mol H}_2\text{SO}_4}{2 \text{ mol NaOH}} = 4.7 \times 10^{-3} \text{ mol H}_2\text{SO}_4$$

$$\text{Concentration of H}_2\text{SO}_4 = \frac{4.7 \times 10^{-3} \text{ mol}}{0.1000 \text{ L}} = 4.7 \times 10^{-2} \text{ M H}_2\text{SO}_4$$



$$0.01420 \text{ L} \times \frac{0.141 \text{ mol H}_3\text{PO}_4}{\text{L}} \times \frac{3 \text{ mol Ba}(\text{OH})_2}{2 \text{ mol H}_3\text{PO}_4} \times \frac{1 \text{ L Ba}(\text{OH})_2}{0.0521 \text{ mol Ba}(\text{OH})_2} = 0.0576 \text{ L}$$

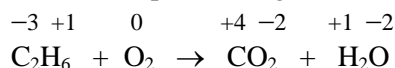
$$= 57.6 \text{ mL Ba}(\text{OH})_2$$

Oxidation-Reduction Reactions

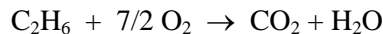
63. a. SrCr_2O_7 : Composed of Sr^{2+} and $\text{Cr}_2\text{O}_7^{2-}$ ions. Sr, +2; O, -2; Cr, $2x + 7(-2) = -2$, $x = +6$
 b. Cu, +2; Cl, -1; c. O, 0; d. H, +1; O, -1
 e. Mg^{2+} and CO_3^{2-} ions present. Mg, +2; O, -2; C, +4; f. Ag, 0
 g. Pb^{2+} and SO_3^{2-} ions present. Pb, +2; O, -2; S, +4; h. O, -2; Pb, +4
 i. Na^+ and $\text{C}_2\text{O}_4^{2-}$ ions present. Na, +1; O, -2; C, $2x + 4(-2) = -2$, $x = +3$
 j. O, -2; C, +4
 k. Ammonium ion has a 1+ charge (NH_4^+), and sulfate ion has a 2- charge (SO_4^{2-}). Therefore, the oxidation state of cerium must be +4 (Ce^{4+}). H, +1; N, -3; O, -2; S, +6
 l. O, -2; Cr, +3
64. a. UO_2^{2+} : O, -2; for U: $x + 2(-2) = +2$, $x = +6$
 b. As_2O_3 : O, -2; for As: $2(x) + 3(-2) = 0$, $x = +3$
 c. NaBiO_3 : Na, +1; O, -2; for Bi: $+1 + x + 3(-2) = 0$, $x = +5$
 d. As_4 : As, 0
 e. HAsO_2 : Assign H = +1 and O = -2; for As: $+1 + x + 2(-2) = 0$, $x = +3$
 f. $\text{Mg}_2\text{P}_2\text{O}_7$: Composed of Mg^{2+} ions and $\text{P}_2\text{O}_7^{4-}$ ions. Mg, +2; O, -2; P, +5

- g. $\text{Na}_2\text{S}_2\text{O}_3$: Composed of Na^+ ions and $\text{S}_2\text{O}_3^{2-}$ ions. Na, +1; O, -2; S, +2
- h. Hg_2Cl_2 : Hg, +1; Cl, -1
- i. $\text{Ca}(\text{NO}_3)_2$: Composed of Ca^{2+} ions and NO_3^- ions. Ca, +2; O, -2; N, +5
65. Apply rules in Table 4.3.
- a. KMnO_4 is composed of K^+ and MnO_4^- ions. Assign oxygen an oxidation state value of -2, which gives manganese a +7 oxidation state because the sum of oxidation states for all atoms in MnO_4^- must equal the 1- charge on MnO_4^- . K, +1; O, -2; Mn, +7.
- b. Assign O a -2 oxidation state, which gives nickel a +4 oxidation state. Ni, +4; O, -2.
- c. $\text{K}_4\text{Fe}(\text{CN})_6$ is composed of K^+ cations and $\text{Fe}(\text{CN})_6^{4-}$ anions. $\text{Fe}(\text{CN})_6^{4-}$ is composed of iron and CN^- anions. For an overall anion charge of 4-, iron must have a +2 oxidation state.
- d. $(\text{NH}_4)_2\text{HPO}_4$ is made of NH_4^+ cations and HPO_4^{2-} anions. Assign +1 as oxidation state of H and -2 as the oxidation state of O. For N in NH_4^+ : $x + 4(+1) = +1$, $x = -3 =$ oxidation state of N. For P in HPO_4^{2-} : $+1 + y + 4(-2) = -2$, $y = +5 =$ oxidation state of P.
- e. O, -2; P, +3
- f. O, -2; Fe, + 8/3
- g. O, -2; F, -1; Xe, +6
- h. F, -1; S, +4
- i. O, -2; C, +2
- j. H, +1; O, -2; C, 0
66. The key to the oxidation states method is to balance the electrons gained by the species reduced with the number of electrons lost from the species oxidized. This is done by assigning oxidation states and, from the change in oxidation states, determining the coefficients necessary to balance electrons gained with electrons lost. After the loss and gain of electrons is balanced, the remainder of the equation is balanced by inspection.
67. a. The species reduced is the element that gains electrons. The reducing agent causes reduction to occur by itself being oxidized. The reducing agent generally refers to the entire formula of the compound/ion that contains the element oxidized.
- b. The species oxidized is the element that loses electrons. The oxidizing agent causes oxidation to occur by itself being reduced. The oxidizing agent generally refers to the entire formula of the compound/ion that contains the element reduced.
- c. For simple binary ionic compounds, the actual charge on the ions are the same as the oxidation states. For covalent compounds and ions, nonzero oxidation states are imaginary charges the elements would have if they were held together by ionic bonds (assuming the bond is between two different nonmetals). Nonzero oxidation states for elements in covalent compounds are not actual charges. Oxidation states for covalent compounds are a bookkeeping method to keep track of electrons in a reaction.

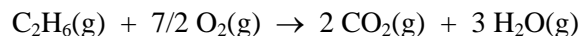
68. a. The first step is to assign oxidation states to all atoms (see numbers above the atoms).



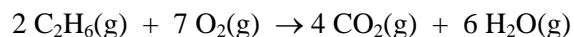
Each carbon atom changes from -3 to $+4$, an increase of seven. Each oxygen atom changes from 0 to -2 , a decrease of 2 . We need $7/2$ O atoms for every C atom.



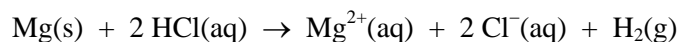
Balancing the remainder of the equation by inspection:



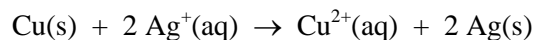
or



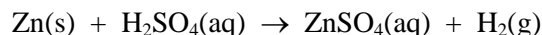
- b. The oxidation state of magnesium changes from 0 to $+2$, an increase of 2 . The oxidation state of hydrogen changes from $+1$ to 0 , a decrease of 1 . We need 2 H atoms for every Mg atom. The balanced equation is:



- c. The oxidation state of copper increases by 2 (0 to $+2$) and the oxidation state of silver decreases by 1 ($+1$ to 0). We need 2 Ag atoms for every Cu atom. The balanced equation is:



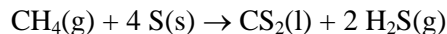
- d. The equation is balanced. Each hydrogen atom gains one electron ($+1 \rightarrow 0$), and each zinc atom loses two electrons ($0 \rightarrow +2$). We need 2 H atoms for every Zn atom. This is the ratio in the given equation:



69. a. $\text{Al}(\text{s}) + 3 \text{HCl}(\text{aq}) \rightarrow \text{AlCl}_3(\text{aq}) + 3/2 \text{H}_2(\text{g})$ or $2 \text{Al}(\text{s}) + 6 \text{HCl}(\text{aq}) \rightarrow 2 \text{AlCl}_3(\text{aq}) + 3 \text{H}_2(\text{g})$

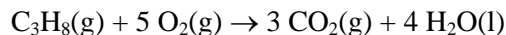
Hydrogen is reduced (goes from the $+1$ oxidation state to the 0 oxidation state), and aluminum Al is oxidized ($0 \rightarrow +3$).

- b. Balancing S is most complicated because sulfur is in both products. Balance C and H first; then worry about S.



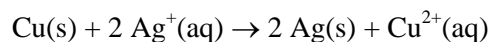
Sulfur is reduced ($0 \rightarrow -2$), and carbon is oxidized ($-4 \rightarrow +4$).

- c. Balance C and H first; then balance O.



Oxygen is reduced ($0 \rightarrow -2$), and carbon is oxidized ($-8/3 \rightarrow +4$).

- d. Although this reaction is mass balanced, it is not charge balanced. We need 2 mol of silver on each side to balance the charge.



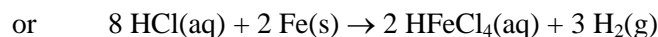
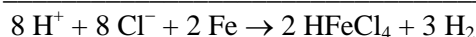
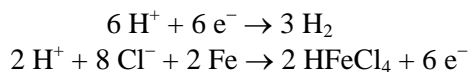
Silver is reduced (+1 \rightarrow 0), and copper is oxidized (0 \rightarrow +2).

70. To determine if the reaction is an oxidation-reduction reaction, assign oxidation states. If the oxidation states change for some elements, then the reaction is a redox reaction. If the oxidation states do not change, then the reaction is not a redox reaction. In redox reactions, the species oxidized (called the reducing agent) shows an increase in the oxidation states, and the species reduced (called the oxidizing agent) shows a decrease in oxidation states.

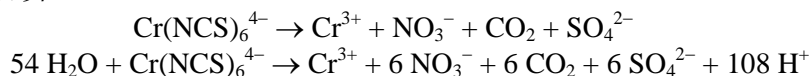
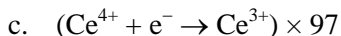
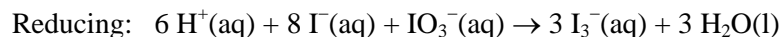
	Redox?	Oxidizing Agent	Reducing Agent	Substance Oxidized	Substance Reduced
a.	Yes	O ₂	CH ₄	CH ₄ (C)	O ₂ (O)
b.	Yes	HCl	Zn	Zn	HCl (H)
c.	No	—	—	—	—
d.	Yes	O ₃	NO	NO (N)	O ₃ (O)
e.	Yes	H ₂ O ₂	H ₂ O ₂	H ₂ O ₂ (O)	H ₂ O ₂ (O)
f.	Yes	CuCl	CuCl	CuCl (Cu)	CuCl (Cu)
g.	No	—	—	—	—
h.	No	—	—	—	—
i.	Yes	SiCl ₄	Mg	Mg	SiCl ₄ (Si)

In c, g, and h, no oxidation states change from reactants to products.

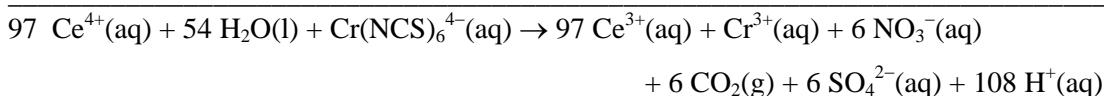
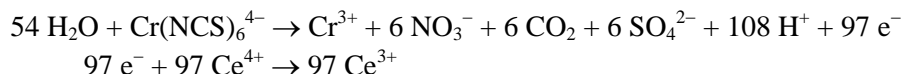
71. a. HCl(aq) dissociates to H⁺(aq) + Cl⁻(aq). For simplicity, let's use H⁺ and Cl⁻ separately.



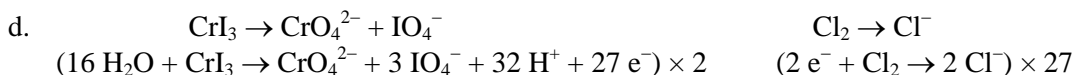
- b.
- $$\begin{array}{l} \text{IO}_3^- \rightarrow \text{I}_3^- \\ 3 \text{IO}_3^- \rightarrow \text{I}_3^- \\ 3 \text{IO}_3^- \rightarrow \text{I}_3^- + 9 \text{H}_2\text{O} \\ 16 \text{e}^- + 18 \text{H}^+ + 3 \text{IO}_3^- \rightarrow \text{I}_3^- + 9 \text{H}_2\text{O} \end{array} \qquad \begin{array}{l} \text{I}^- \rightarrow \text{I}_3^- \\ (3 \text{I}^- \rightarrow \text{I}_3^- + 2 \text{e}^-) \times 8 \end{array}$$
- $$\begin{array}{l} 16 \text{e}^- + 18 \text{H}^+ + 3 \text{IO}_3^- \rightarrow \text{I}_3^- + 9 \text{H}_2\text{O} \\ 24 \text{I}^- \rightarrow 8 \text{I}_3^- + 16 \text{e}^- \end{array}$$
-
- $$18 \text{H}^+ + 24 \text{I}^- + 3 \text{IO}_3^- \rightarrow 9 \text{I}_3^- + 9 \text{H}_2\text{O}$$



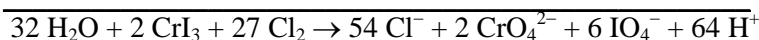
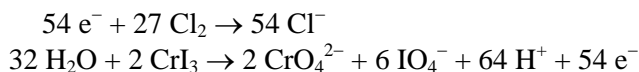
Charge on left = -4. Charge on right = +3 + 6(-1) + 6(-2) + 108(+1) = +93. Add 97 e⁻ to the product side, and then add the two balanced half-reactions with a common factor of 97 e⁻ transferred.



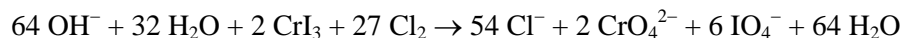
This is very complicated. A check of the net charge is a good check to see if the equation is balanced. Left: charge = 97(+4) - 4 = +384. Right: charge = 97(+3) + 3 + 6(-1) + 6(-2) + 108(+1) = +384.



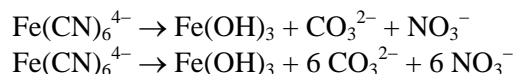
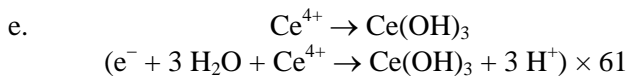
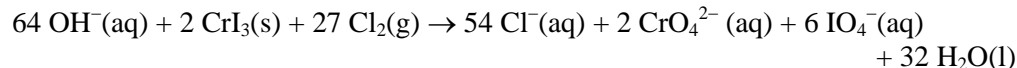
Common factor is a transfer of 54 e⁻.



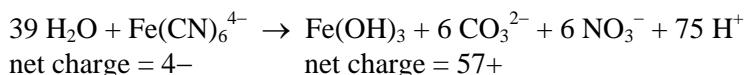
Add 64 OH⁻ to both sides and convert 64 H⁺ into 64 H₂O.



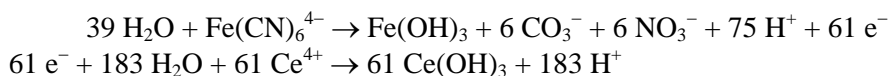
Reducing gives:



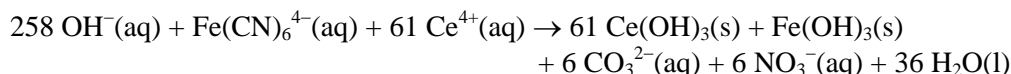
There are 39 extra O atoms on right. Add 39 H₂O to left, then add 75 H⁺ to right to balance H⁺.



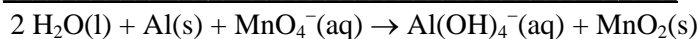
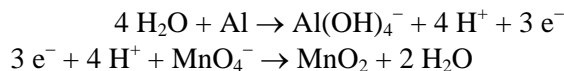
Add $61 e^-$ to the product side, and then add the two balanced half-reactions with a common factor of $61 e^-$ transferred.



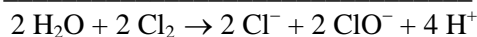
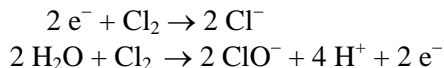
Adding 258OH^- to each side, and then reducing gives:



72. Use the same method as with acidic solutions. After the final balanced equation, convert H^+ to OH^- as described in section 4.11 of the text. The extra step involves converting H^+ into H_2O by adding equal moles of OH^- to each side of the reaction. This converts the reaction to a basic solution while keeping it balanced.



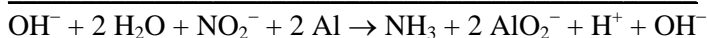
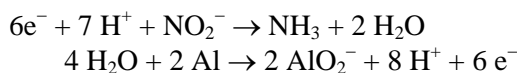
Because H^+ doesn't appear in the final balanced reaction, we are done.



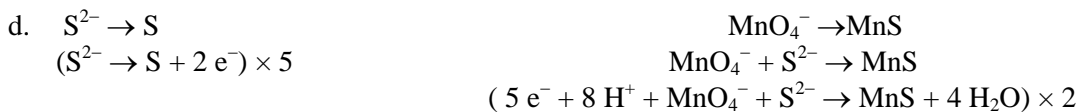
Now convert to a basic solution. Add 4OH^- to both sides of the equation. The 4OH^- will react with the 4H^+ on the product side to give $4 \text{H}_2\text{O}$. After this step, cancel identical species on both sides ($2 \text{H}_2\text{O}$). Applying these steps gives $4 \text{OH}^- + 2 \text{Cl}_2 \rightarrow 2 \text{Cl}^- + 2 \text{ClO}^- + 2 \text{H}_2\text{O}$, which can be further simplified to:



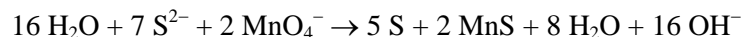
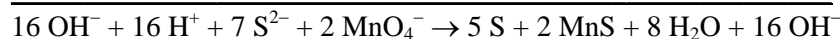
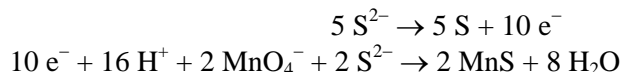
Common factor is a transfer of $6 e^-$.



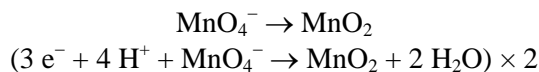
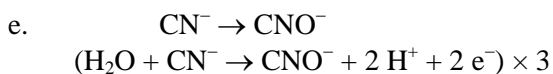
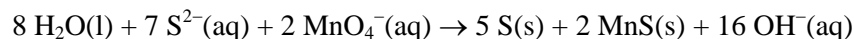
Reducing gives: $OH^-(aq) + H_2O(l) + NO_2^-(aq) + 2Al(s) \rightarrow NH_3(g) + 2AlO_2^-(aq)$



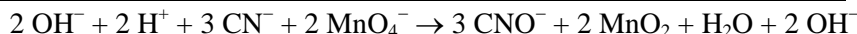
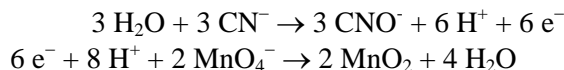
Common factor is a transfer of $10e^-$.



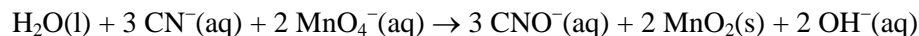
Reducing gives:



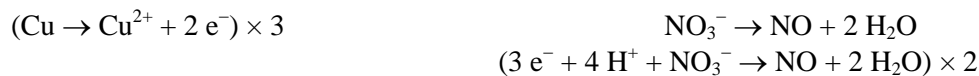
Common factor is a transfer of 6 electrons.



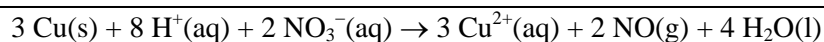
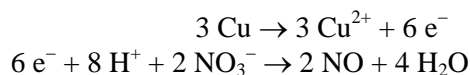
Reducing gives:



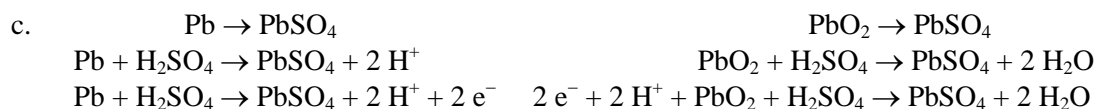
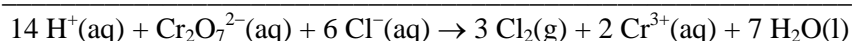
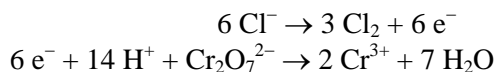
73. a. Review Section 4.11 of the text for rules on balancing by the half-reaction method. The first step is to separate the reaction into two half-reactions, and then balance each half-reaction separately.



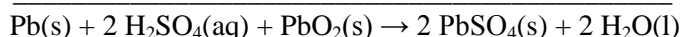
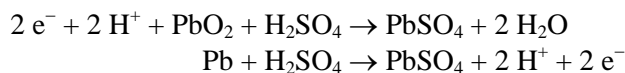
Adding the two balanced half-reactions so electrons cancel:



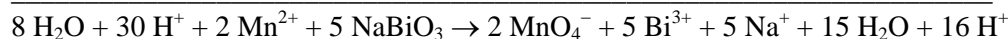
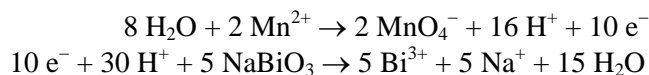
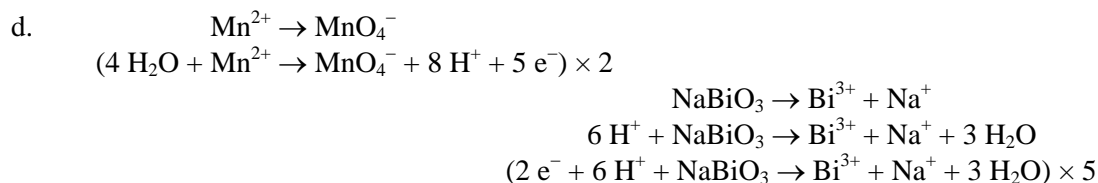
Add the two balanced half-reactions with six electrons transferred:



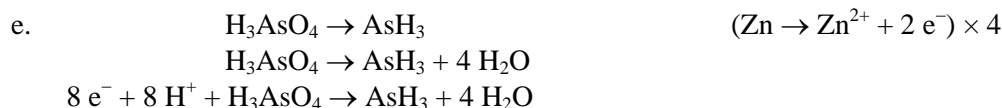
Add the two half-reactions with two electrons transferred:

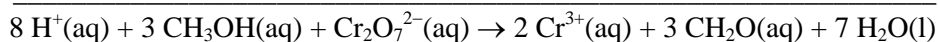
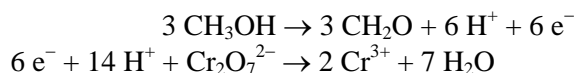
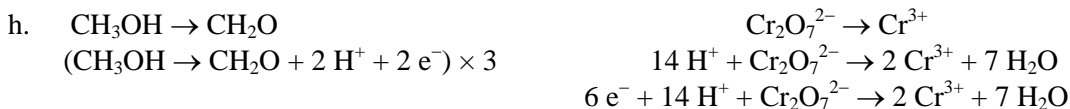
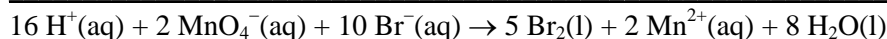
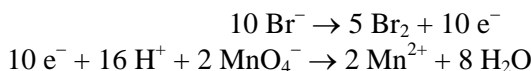
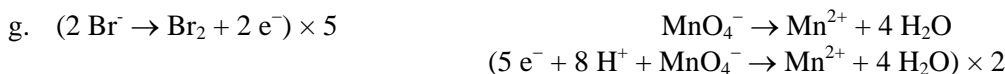
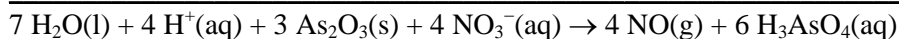
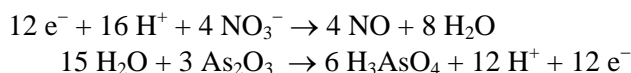
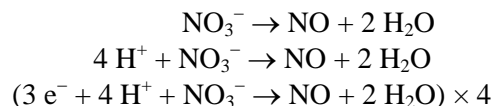
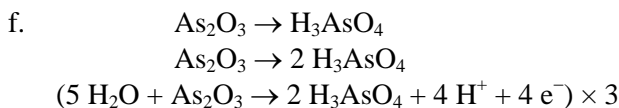
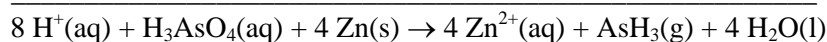
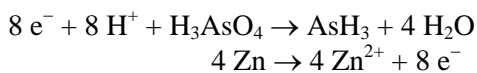


This is the reaction that occurs in an automobile lead storage battery.



Simplifying:

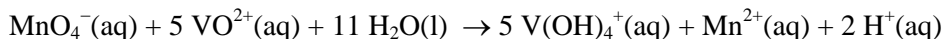




74. The unbalanced reaction is: $VO^{2+} + MnO_4^{-} \rightarrow V(OH)_4^{+} + Mn^{2+}$

This is a redox reaction in acidic solution and must be balanced accordingly. The two half-reactions to balance are: $VO^{2+} \rightarrow V(OH)_4^{+}$ and $MnO_4^{-} \rightarrow Mn^{2+}$

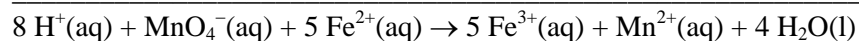
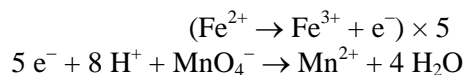
Balancing by the half-reaction method gives:



$$0.02645 \text{ L} \times \frac{0.02250 \text{ mol MnO}_4^{-}}{\text{L}} \times \frac{5 \text{ mol VO}^{2+}}{\text{mol MnO}_4^{-}} \times \frac{1 \text{ mol V}}{\text{mol VO}^{2+}} \times \frac{50.94 \text{ g V}}{\text{mol V}} = 0.1516 \text{ g V}$$

$$0.581 = \frac{0.1516 \text{ g V}}{\text{mass of ore sample}}, \quad 0.1516/0.581 = 0.261 \text{ g ore sample}$$

75.



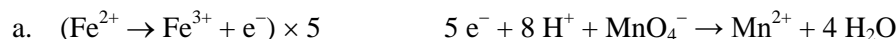
From the titration data we can get the number of moles of Fe^{2+} . We then convert this to a mass of iron and calculate the mass percent of iron in the sample.

$$\begin{aligned} 38.37 \times 10^{-3} \text{ L MnO}_4^- \times \frac{0.0198 \text{ mol MnO}_4^-}{\text{L}} \times \frac{5 \text{ mol Fe}^{2+}}{\text{mol MnO}_4^-} &= 3.80 \times 10^{-3} \text{ mol Fe}^{2+} \\ &= 3.80 \times 10^{-3} \text{ mol Fe present} \end{aligned}$$

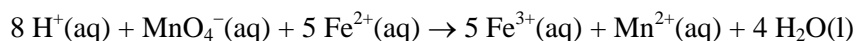
$$3.80 \times 10^{-3} \text{ mol Fe} \times \frac{55.85 \text{ g Fe}}{\text{mol Fe}} = 0.212 \text{ g Fe}$$

$$\text{Mass \% Fe} = \frac{0.212 \text{ g}}{0.6128 \text{ g}} \times 100 = 34.6\% \text{ Fe}$$

76.

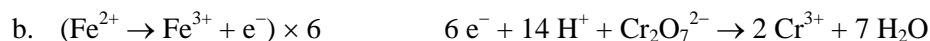


The balanced equation is:

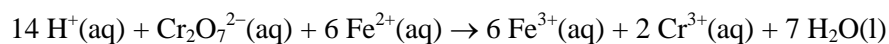


$$20.62 \times 10^{-3} \text{ L soln} \times \frac{0.0216 \text{ mol MnO}_4^-}{\text{L soln}} \times \frac{5 \text{ mol Fe}^{2+}}{\text{mol MnO}_4^-} = 2.23 \times 10^{-3} \text{ mol Fe}^{2+}$$

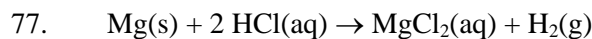
$$\text{Molarity} = \frac{2.23 \times 10^{-3} \text{ mol Fe}^{2+}}{50.00 \times 10^{-3} \text{ L}} = 4.46 \times 10^{-2} \text{ M Fe}^{2+}$$



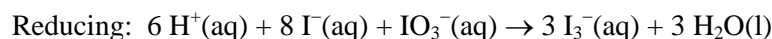
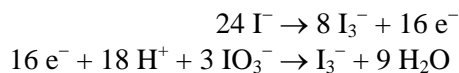
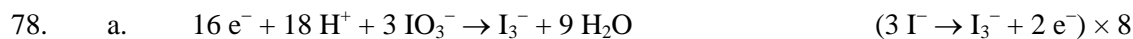
The balanced equation is:



$$\begin{aligned} 50.00 \times 10^{-3} \text{ L} \times \frac{4.46 \times 10^{-2} \text{ mol Fe}^{2+}}{\text{L}} \times \frac{1 \text{ mol Cr}_2\text{O}_7^{2-}}{6 \text{ mol Fe}^{2+}} \times \frac{1 \text{ L}}{0.0150 \text{ mol Cr}_2\text{O}_7^{2-}} \\ = 2.48 \times 10^{-2} \text{ L or } 24.8 \text{ mL} \end{aligned}$$



$$3.00 \text{ g Mg} \times \frac{1 \text{ mol Mg}}{24.31 \text{ g Mg}} \times \frac{2 \text{ mol HCl}}{\text{mol Mg}} \times \frac{1 \text{ L HCl}}{5.0 \text{ mol HCl}} = 0.0494 \text{ L} = 49.4 \text{ mL HCl}$$



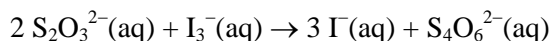
b. $0.6013 \text{ g KIO}_3 \times \frac{1 \text{ mol KIO}_3}{214.0 \text{ g KIO}_3} = 2.810 \times 10^{-3} \text{ mol KIO}_3$

$$2.810 \times 10^{-3} \text{ mol KIO}_3 \times \frac{8 \text{ mol KI}}{\text{mol KIO}_3} \times \frac{166.0 \text{ g KI}}{\text{mol KI}} = 3.732 \text{ g KI}$$

$$2.810 \times 10^{-3} \text{ mol KIO}_3 \times \frac{6 \text{ mol HCl}}{\text{mol KIO}_3} \times \frac{1 \text{ L}}{3.00 \text{ mol HCl}} = 5.62 \times 10^{-3} \text{ L} = 5.62 \text{ mL HCl}$$



Adding the balanced half-reactions gives:

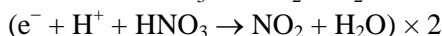
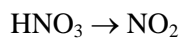


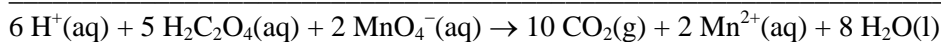
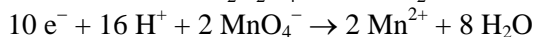
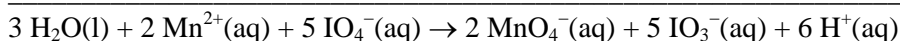
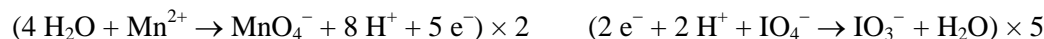
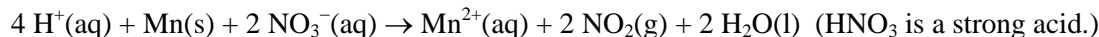
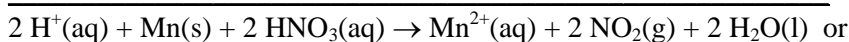
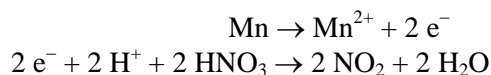
d. $25.00 \times 10^{-3} \text{ L KIO}_3 \times \frac{0.0100 \text{ mol KIO}_3}{\text{L}} \times \frac{3 \text{ mol I}_3^-}{\text{mol KIO}_3} \times \frac{2 \text{ mol Na}_2\text{S}_2\text{O}_3}{\text{mol I}_3^-} =$
 $1.50 \times 10^{-3} \text{ mol Na}_2\text{S}_2\text{O}_3$

$$M_{\text{Na}_2\text{S}_2\text{O}_3} = \frac{1.50 \times 10^{-3} \text{ mol}}{32.04 \times 10^{-3} \text{ L}} = 0.0468 \text{ M Na}_2\text{S}_2\text{O}_3$$

e. $0.5000 \text{ L} \times \frac{0.0100 \text{ mol KIO}_3}{\text{L}} \times \frac{214.0 \text{ g KIO}_3}{\text{mol KIO}_3} = 1.07 \text{ g KIO}_3$

Place 1.07 g KIO_3 in a 500-mL volumetric flask; add water to dissolve the KIO_3 ; continue adding water to the 500.0-mL mark, with mixing along the way.





$$0.1058 \text{ g H}_2\text{C}_2\text{O}_4 \times \frac{1 \text{ mol H}_2\text{C}_2\text{O}_4}{90.034 \text{ g}} \times \frac{2 \text{ mol MnO}_4^{-}}{5 \text{ mol H}_2\text{C}_2\text{O}_4} = 4.700 \times 10^{-4} \text{ mol MnO}_4^{-}$$

$$\text{Molarity} = \frac{4.700 \times 10^{-4} \text{ mol MnO}_4^{-}}{28.97 \text{ mL}} \times \frac{1000 \text{ mL}}{\text{L}} = 1.622 \times 10^{-2} \text{ M MnO}_4^{-}$$

Additional Exercises

81. Mol of KHP used = $0.4016 \text{ g} \times \frac{1 \text{ mol}}{204.22 \text{ g}} = 1.967 \times 10^{-3} \text{ mol KHP}$

Because 1 mole of NaOH reacts completely with 1 mole of KHP, the NaOH solution contains $1.967 \times 10^{-3} \text{ mol NaOH}$.

$$\text{Molarity of NaOH} = \frac{1.967 \times 10^{-3} \text{ mol}}{25.06 \times 10^{-3} \text{ L}} = \frac{7.849 \times 10^{-2} \text{ mol NaOH}}{\text{L}}$$

$$\text{Maximum molarity} = \frac{1.967 \times 10^{-3} \text{ mol}}{25.01 \times 10^{-3} \text{ L}} = \frac{7.865 \times 10^{-2} \text{ mol NaOH}}{\text{L}}$$

$$\text{Minimum molarity} = \frac{1.967 \times 10^{-3} \text{ mol}}{25.11 \times 10^{-3} \text{ L}} = \frac{7.834 \times 10^{-2} \text{ mol NaOH}}{\text{L}}$$

We can express this as $0.07849 \pm 0.00016 M$. An alternate way is to express the molarity as $0.0785 \pm 0.0002 M$. This second way shows the actual number of significant figures in the molarity. The advantage of the first method is that it shows that we made all our individual measurements to four significant figures.

82. Desired uncertainty is 1% of 0.02, or ± 0.0002 . So we want the solution to be $0.0200 \pm 0.0002 M$, or the concentration should be between 0.0198 and 0.0202 M . We should use a 1-L volumetric flask to make the solution. They are good to $\pm 0.1\%$. We want to weigh out between 0.0198 mol and 0.0202 mol of KIO_3 .

Molar mass of $\text{KIO}_3 = 39.10 + 126.9 + 3(16.00) = 214.0 \text{ g/mol}$

$$0.0198 \text{ mol} \times \frac{214.0 \text{ g}}{\text{mol}} = 4.237 \text{ g}; \quad 0.0202 \text{ mol} \times \frac{214.0 \text{ g}}{\text{mol}} = 4.323 \text{ g} \text{ (carrying extra sig. figs.)}$$

We should weigh out between 4.24 and 4.32 g of KIO_3 . We should weigh it to the nearest milligram or 0.1 mg. Dissolve the KIO_3 in water, and dilute to the mark in a 1-liter volumetric flask. This will produce a solution whose concentration is within the limits and is known to at least the fourth decimal place.

83. $\text{H}_2\text{SO}_4(\text{aq}) + 2 \text{NaOH}(\text{aq}) \rightarrow \text{Na}_2\text{SO}_4(\text{aq}) + 2 \text{H}_2\text{O}(\text{l})$

$$0.02844 \text{ L} \times \frac{0.1000 \text{ mol NaOH}}{\text{L}} \times \frac{1 \text{ mol H}_2\text{SO}_4}{2 \text{ mol NaOH}} \times \frac{1 \text{ mol SO}_2}{\text{mol H}_2\text{SO}_4} \times \frac{32.07 \text{ g S}}{\text{mol SO}_2} = 4.560 \times 10^{-2} \text{ g S}$$

$$\text{Mass \% S} = \frac{0.04560 \text{ g}}{1.325 \text{ g}} \times 100 = 3.442\%$$

84. $3 (\text{NH}_4)_2\text{CrO}_4(\text{aq}) + 2 \text{Cr}(\text{NO}_2)_3(\text{aq}) \rightarrow 6 \text{NH}_4\text{NO}_2(\text{aq}) + \text{Cr}_2(\text{CrO}_4)_3(\text{s})$

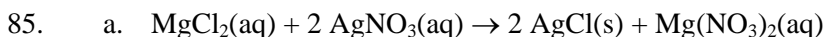
$$0.203 \text{ L} \times \frac{0.307 \text{ mol}}{\text{L}} = 6.23 \times 10^{-2} \text{ mol } (\text{NH}_4)_2\text{CrO}_4$$

$$0.137 \text{ L} \times \frac{0.269 \text{ mol}}{\text{L}} = 3.69 \times 10^{-2} \text{ mol } \text{Cr}(\text{NO}_2)_3$$

$0.0623 \text{ mol} / 0.0369 \text{ mol} = 1.69$ (actual); the balanced reaction requires a $3/2 = 1.5$ to 1 mole ratio between $(\text{NH}_4)_2\text{CrO}_4$ and $\text{Cr}(\text{NO}_2)_3$. Actual > required, so $\text{Cr}(\text{NO}_2)_3$ (the denominator) is limiting.

$$3.69 \times 10^{-2} \text{ mol } \text{Cr}(\text{NO}_2)_3 \times \frac{1 \text{ mol } \text{Cr}_2(\text{CrO}_4)_3}{2 \text{ mol } \text{Cr}(\text{NO}_2)_3} \times \frac{452.00 \text{ g } \text{Cr}_2(\text{CrO}_4)_3}{\text{mol } \text{Cr}_2(\text{CrO}_4)_3} = 8.34 \text{ g } \text{Cr}_2(\text{CrO}_4)_3$$

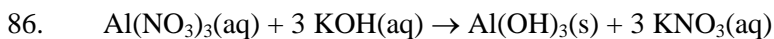
$$0.880 = \frac{\text{actual yield}}{8.34 \text{ g}}, \quad \text{actual yield} = (8.34 \text{ g})(0.880) = 7.34 \text{ g } \text{Cr}_2(\text{CrO}_4)_3 \text{ isolated}$$



$$0.641 \text{ g AgCl} \times \frac{1 \text{ mol AgCl}}{143.4 \text{ g AgCl}} \times \frac{1 \text{ mol MgCl}_2}{2 \text{ mol AgCl}} \times \frac{95.21 \text{ g}}{\text{mol MgCl}_2} = 0.213 \text{ g MgCl}_2$$

$$\frac{0.213 \text{ g MgCl}_2}{1.50 \text{ g mixture}} \times 100 = 14.2\% \text{ MgCl}_2$$

b. $0.213 \text{ g MgCl}_2 \times \frac{1 \text{ mol MgCl}_2}{95.21 \text{ g}} \times \frac{2 \text{ mol AgNO}_3}{\text{mol MgCl}_2} \times \frac{1 \text{ L}}{0.500 \text{ mol AgNO}_3} \times \frac{1000 \text{ mL}}{1 \text{ L}}$
 $= 8.95 \text{ mL AgNO}_3$

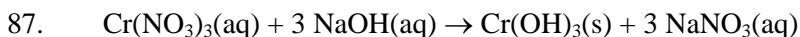


$$0.0500 \text{ L} \times \frac{0.200 \text{ mol Al}(\text{NO}_3)_3}{\text{L}} = 0.0100 \text{ mol Al}(\text{NO}_3)_3$$

$$0.2000 \text{ L} \times \frac{0.100 \text{ mol KOH}}{\text{L}} = 0.0200 \text{ mol KOH}$$

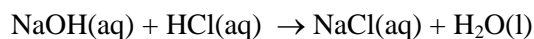
From the balanced equation, 3 moles of KOH are required to react with 1 mole of $\text{Al}(\text{NO}_3)_3$ (3 : 1 mole ratio). The actual KOH to $\text{Al}(\text{NO}_3)_3$ mole ratio present is $0.0200/0.0100 = 2$ (2 : 1). Because the actual mole ratio present is less than the required mole ratio, KOH is the limiting reagent.

$$0.0200 \text{ mol KOH} \times \frac{1 \text{ mol Al}(\text{OH})_3}{3 \text{ mol KOH}} \times \frac{78.00 \text{ g Al}(\text{OH})_3}{\text{mol Al}(\text{OH})_3} = 0.520 \text{ g Al}(\text{OH})_3$$



$$\text{Mol NaOH used} = 2.06 \text{ g Cr}(\text{OH})_3 \times \frac{1 \text{ mol Cr}(\text{OH})_3}{103.02 \text{ g}} \times \frac{3 \text{ mol NaOH}}{\text{mol Cr}(\text{OH})_3} = 6.00 \times 10^{-2} \text{ mol}$$

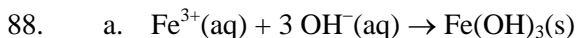
to form precipitate



$$\text{Mol NaOH used} = 0.1000 \text{ L} \times \frac{0.400 \text{ mol HCl}}{\text{L}} \times \frac{1 \text{ mol NaOH}}{\text{mol HCl}} = 4.00 \times 10^{-2} \text{ mol}$$

to react with HCl

$$M_{\text{NaOH}} = \frac{\text{total mol NaOH}}{\text{volume}} = \frac{6.00 \times 10^{-2} \text{ mol} + 4.00 \times 10^{-2} \text{ mol}}{0.0500 \text{ L}} = 2.00 \text{ M NaOH}$$



$$\text{Fe}(\text{OH})_3: 55.85 + 3(16.00) + 3(1.008) = 106.87 \text{ g/mol}$$

$$0.107 \text{ g Fe(OH)}_3 \times \frac{55.85 \text{ g Fe}}{106.9 \text{ g Fe(OH)}_3} = 0.0559 \text{ g Fe}$$

b. $\text{Fe(NO}_3)_3$: $55.85 + 3(14.01) + 9(16.00) = 241.86 \text{ g/mol}$

$$0.0559 \text{ g Fe} \times \frac{241.9 \text{ g Fe(NO}_3)_3}{55.85 \text{ g Fe}} = 0.242 \text{ g Fe(NO}_3)_3$$

c. $\text{Mass \% Fe(NO}_3)_3 = \frac{0.242 \text{ g}}{0.456 \text{ g}} \times 100 = 53.1\%$

89. $\text{Mol CaCl}_2 \text{ present} = 0.230 \text{ L CaCl}_2 \times \frac{0.275 \text{ mol CaCl}_2}{\text{L CaCl}_2} = 6.33 \times 10^{-2} \text{ mol CaCl}_2$

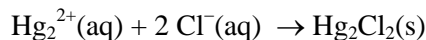
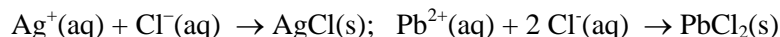
The volume of CaCl_2 solution after evaporation is:

$$6.33 \times 10^{-2} \text{ mol CaCl}_2 \times \frac{1 \text{ L CaCl}_2}{1.10 \text{ mol CaCl}_2} = 5.75 \times 10^{-2} \text{ L} = 57.5 \text{ mL CaCl}_2$$

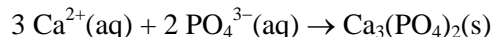
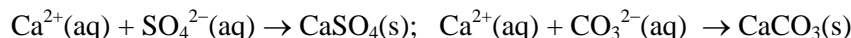
Volume H_2O evaporated = $230. \text{ mL} - 57.5 \text{ mL} = 173 \text{ mL H}_2\text{O}$ evaporated

90. There are other possible correct choices for the following answers. We have listed only three possible reactants in each case.

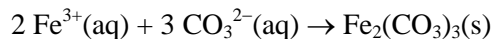
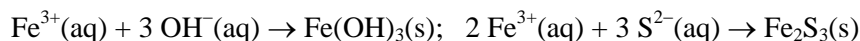
a. AgNO_3 , $\text{Pb(NO}_3)_2$, and $\text{Hg}_2(\text{NO}_3)_2$ would form precipitates with the Cl^- ion.



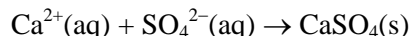
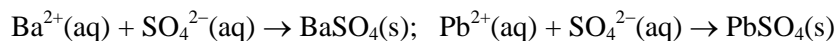
b. Na_2SO_4 , Na_2CO_3 , and Na_3PO_4 would form precipitates with the Ca^{2+} ion.



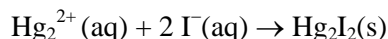
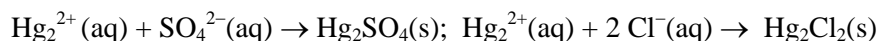
c. NaOH , Na_2S , and Na_2CO_3 would form precipitates with the Fe^{3+} ion.



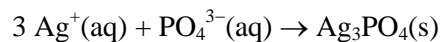
d. BaCl_2 , $\text{Pb(NO}_3)_2$, and $\text{Ca(NO}_3)_2$ would form precipitates with the SO_4^{2-} ion.



e. Na_2SO_4 , NaCl , and NaI would form precipitates with the Hg_2^{2+} ion.



f. NaBr, Na₂CrO₄, and Na₃PO₄ would form precipitates with the Ag⁺ ion.



$$91. \quad \text{a.} \quad 0.308 \text{ g AgCl} \times \frac{35.45 \text{ g Cl}}{143.4 \text{ g AgCl}} = 0.0761 \text{ g Cl}; \quad \% \text{ Cl} = \frac{0.0761 \text{ g}}{0.256 \text{ g}} \times 100 = 29.7\% \text{ Cl}$$

Cobalt(III) oxide, Co₂O₃: 2(58.93) + 3(16.00) = 165.86 g/mol

$$0.145 \text{ g Co}_2\text{O}_3 \times \frac{117.86 \text{ g Co}}{165.86 \text{ g Co}_2\text{O}_3} = 0.103 \text{ g Co}; \quad \% \text{ Co} = \frac{0.103 \text{ g}}{0.416 \text{ g}} \times 100 = 24.8\% \text{ Co}$$

The remainder, 100.0 – (29.7 + 24.8) = 45.5%, is water.

Assuming 100.0 g of compound:

$$45.5 \text{ g H}_2\text{O} \times \frac{2.016 \text{ g H}}{18.02 \text{ g H}_2\text{O}} = 5.09 \text{ g H}; \quad \% \text{ H} = \frac{5.09 \text{ g H}}{100.0 \text{ g compound}} \times 100 = 5.09\% \text{ H}$$

$$45.5 \text{ g H}_2\text{O} \times \frac{16.00 \text{ g O}}{18.02 \text{ g H}_2\text{O}} = 40.4 \text{ g O}; \quad \% \text{ O} = \frac{40.4 \text{ g O}}{100.0 \text{ g compound}} \times 100 = 40.4\% \text{ O}$$

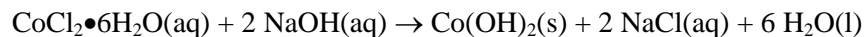
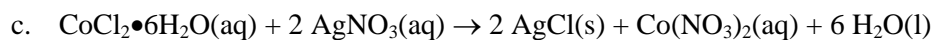
The mass percent composition is 24.8% Co, 29.7% Cl, 5.09% H, and 40.4% O.

b. Out of 100.0 g of compound, there are:

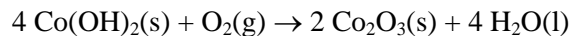
$$24.8 \text{ g Co} \times \frac{1 \text{ mol}}{58.93 \text{ g Co}} = 0.421 \text{ mol Co}; \quad 29.7 \text{ g Cl} \times \frac{1 \text{ mol}}{35.45 \text{ g Cl}} = 0.838 \text{ mol Cl}$$

$$5.09 \text{ g H} \times \frac{1 \text{ mol}}{1.008 \text{ g H}} = 5.05 \text{ mol H}; \quad 40.4 \text{ g O} \times \frac{1 \text{ mol}}{16.00 \text{ g O}} = 2.53 \text{ mol O}$$

Dividing all results by 0.421, we get CoCl₂•6H₂O for the empirical formula, which is also the molecular formula.



Co(OH)₂ → Co₂O₃ This is an oxidation-reduction reaction. Thus we also need to include an oxidizing agent. The obvious choice is O₂.



92. a. Assume 100.00 g of material.

$$42.23 \text{ g C} \times \frac{1 \text{ mol C}}{12.011 \text{ g C}} = 3.516 \text{ mol C}; \quad 55.66 \text{ g F} \times \frac{1 \text{ mol F}}{19.00 \text{ g F}} = 2.929 \text{ mol F}$$

$$2.11 \text{ g B} \times \frac{1 \text{ mol B}}{10.81 \text{ g B}} = 0.195 \text{ mol B}$$

$$\text{Dividing by the smallest number: } \frac{3.516}{0.195} = 18.0; \quad \frac{2.929}{0.195} = 15.0$$

The empirical formula is $\text{C}_{18}\text{F}_{15}\text{B}$.

b. $0.3470 \text{ L} \times \frac{0.01267 \text{ mol}}{\text{L}} = 4.396 \times 10^{-3} \text{ mol BARF}$

$$\text{Molar mass of BARF} = \frac{2.251 \text{ g}}{4.396 \times 10^{-3} \text{ mol}} = 512.1 \text{ g/mol}$$

The empirical formula mass of BARF is 511.99 g. Therefore, the molecular formula is the same as the empirical formula, $\text{C}_{18}\text{F}_{15}\text{B}$.

93. $\text{Ag}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \rightarrow \text{AgCl}(\text{s})$; let $x = \text{mol NaCl}$ and $y = \text{mol KCl}$.

$$(22.90 \times 10^{-3} \text{ L}) \times 0.1000 \text{ mol/L} = 2.290 \times 10^{-3} \text{ mol Ag}^+ = 2.290 \times 10^{-3} \text{ mol Cl}^- \text{ total}$$

$$x + y = 2.290 \times 10^{-3} \text{ mol Cl}^-, \quad x = 2.290 \times 10^{-3} - y$$

Because the molar mass of NaCl is 58.44 g/mol and the molar mass of KCl is 74.55 g/mol:

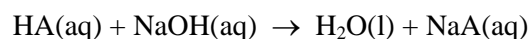
$$(58.44)x + (74.55)y = 0.1586 \text{ g}$$

$$58.44(2.290 \times 10^{-3} - y) + (74.55)y = 0.1586, \quad (16.11)y = 0.0248, \quad y = 1.54 \times 10^{-3} \text{ mol KCl}$$

$$\text{Mass \% KCl} = \frac{1.54 \times 10^{-3} \text{ mol} \times 74.55 \text{ g/mol}}{0.1586 \text{ g}} \times 100 = 72.4\% \text{ KCl}$$

$$\% \text{ NaCl} = 100.0 - 72.4 = 27.6\% \text{ NaCl}$$

94. Using HA as an abbreviation for the monoprotic acid acetylsalicylic acid:



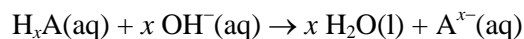
$$\text{Mol HA} = 0.03517 \text{ L NaOH} \times \frac{0.5065 \text{ mol NaOH}}{\text{L NaOH}} \times \frac{1 \text{ mol HA}}{\text{mol NaOH}} = 1.781 \times 10^{-2} \text{ mol HA}$$

From the problem, 3.210 g HA was reacted, so:

$$\text{molar mass} = \frac{3.210 \text{ g HA}}{1.781 \times 10^{-2} \text{ mol HA}} = 180.2 \text{ g/mol}$$

$$95. \quad \text{Mol C}_6\text{H}_8\text{O}_7 = 0.250 \text{ g C}_6\text{H}_8\text{O}_7 \times \frac{1 \text{ mol C}_6\text{H}_8\text{O}_7}{192.1 \text{ g C}_6\text{H}_8\text{O}_7} = 1.30 \times 10^{-3} \text{ mol C}_6\text{H}_8\text{O}_7$$

Let H_xA represent citric acid, where x is the number of acidic hydrogens. The balanced neutralization reaction is:



$$\text{Mol OH}^- \text{ reacted} = 0.0372 \text{ L} \times \frac{0.105 \text{ mol OH}^-}{\text{L}} = 3.91 \times 10^{-3} \text{ mol OH}^-$$

$$x = \frac{\text{mol OH}^-}{\text{mol citric acid}} = \frac{3.91 \times 10^{-3} \text{ mol}}{1.30 \times 10^{-3} \text{ mol}} = 3.01$$

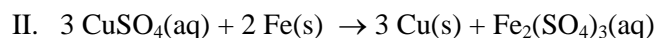
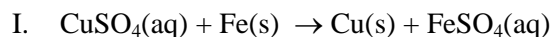
Therefore, the general acid formula for citric acid is H_3A , meaning that citric acid has three acidic hydrogens per citric acid molecule (citric acid is a triprotic acid).

Challenge Problems

$$96. \quad \text{Mol CuSO}_4 = 87.6 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.500 \text{ mol}}{\text{L}} = 0.0439 \text{ mol}$$

$$\text{Mol Fe} = 2.00 \text{ g} \times \frac{1 \text{ mol Fe}}{55.85 \text{ g}} = 0.0358 \text{ mol}$$

The two possible reactions are:



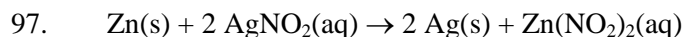
If reaction I occurs, Fe is limiting, and we can produce:

$$0.0358 \text{ mol Fe} \times \frac{1 \text{ mol Cu}}{1 \text{ mol Fe}} \times \frac{63.55 \text{ g Cu}}{\text{mol Cu}} = 2.28 \text{ g Cu}$$

If reaction II occurs, CuSO_4 is limiting, and we can produce:

$$0.0439 \text{ mol CuSO}_4 \times \frac{3 \text{ mol Cu}}{3 \text{ mol CuSO}_4} \times \frac{63.55 \text{ g Cu}}{\text{mol Cu}} = 2.79 \text{ g Cu}$$

Assuming 100% yield, reaction I occurs because it fits the data best.



Let x = mass of Ag and y = mass of Zn after the reaction has stopped. Then $x + y = 29.0$ g. Because the moles of Ag produced will equal two times the moles of Zn reacted:

$$(19.0 - y) \text{ g Zn} \times \frac{1 \text{ mol Zn}}{65.38 \text{ g Zn}} \times \frac{2 \text{ mol Ag}}{1 \text{ mol Zn}} = x \text{ g Ag} \times \frac{1 \text{ mol Ag}}{107.9 \text{ g Ag}}$$

Simplifying:

$$3.059 \times 10^{-2}(19.0 - y) = (9.268 \times 10^{-3})x$$

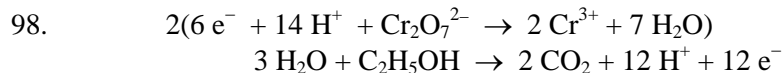
Substituting $x = 29.0 - y$ into the equation gives:

$$3.059 \times 10^{-2}(19.0 - y) = 9.268 \times 10^{-3}(29.0 - y)$$

Solving:

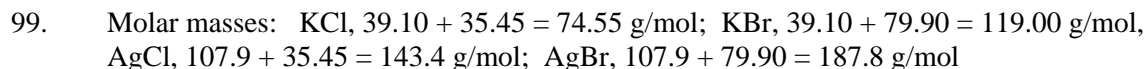
$$0.581 - (3.059 \times 10^{-2})y = 0.269 - (9.268 \times 10^{-3})y, (2.132 \times 10^{-2})y = 0.312, y = 14.6 \text{ g Zn}$$

14.6 g Zn are present, and $29.0 - 14.6 = 14.4$ g Ag are also present after the reaction is stopped.



$$0.03105 \text{ L} \left(\frac{0.0600 \text{ mol Cr}_2\text{O}_7^{2-}}{\text{L}} \right) \left(\frac{1 \text{ mol C}_2\text{H}_5\text{OH}}{2 \text{ mol Cr}_2\text{O}_7^{2-}} \right) \left(\frac{46.07 \text{ g}}{\text{mol C}_2\text{H}_5\text{OH}} \right) = 0.0429 \text{ g C}_2\text{H}_5\text{OH}$$

$$\frac{0.0429 \text{ g C}_2\text{H}_5\text{OH}}{30.0 \text{ g blood}} \times 100 = 0.143\% \text{ C}_2\text{H}_5\text{OH}$$



Let x = number of moles of KCl in mixture and y = number of moles of KBr in mixture. $\text{Ag}^+ + \text{Cl}^- \rightarrow \text{AgCl}$ and $\text{Ag}^+ + \text{Br}^- \rightarrow \text{AgBr}$; so, x = moles AgCl and y = moles AgBr.

Setting up two equations from the given information:

$$0.1024 \text{ g} = (74.55)x + (119.0)y \quad \text{and} \quad 0.1889 \text{ g} = (143.4)x + (187.8)y$$

Multiply the first equation by $\frac{187.8}{119.0}$, and then subtract from the second.

$$\begin{aligned} 0.1889 &= (143.4)x + (187.8)y \\ -0.1616 &= -(117.7)x - (187.8)y \\ \hline 0.0273 &= (25.7)x, \quad x = 1.06 \times 10^{-3} \text{ mol KCl} \end{aligned}$$

$$1.06 \times 10^{-3} \text{ mol KCl} \times \frac{74.55 \text{ g KCl}}{\text{mol KCl}} = 0.0790 \text{ g KCl}$$

$$\text{Mass \% KCl} = \frac{0.0790 \text{ g}}{0.1024 \text{ g}} \times 100 = 77.1\%, \quad \% \text{ KBr} = 100.0 - 77.1 = 22.9\%$$

100. Let x = mass of NaCl, and let y = mass K_2SO_4 . So $x + y = 10.00$.

Two reactions occur: $\text{Pb}^{2+}(\text{aq}) + 2 \text{Cl}^{-}(\text{aq}) \rightarrow \text{PbCl}_2(\text{s})$ and
 $\text{Pb}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{PbSO}_4(\text{s})$

Molar mass of NaCl = 58.44 g/mol; molar mass of K_2SO_4 = 174.27 g/mol; molar mass of PbCl_2 = 278.1 g/mol; molar mass of PbSO_4 = 303.3 g/mol

$$\frac{x}{58.44} = \text{moles NaCl}; \quad \frac{y}{174.27} = \text{moles K}_2\text{SO}_4$$

$$\begin{aligned} \text{mass of PbCl}_2 + \text{mass PbSO}_4 &= \text{total mass of solid} \\ \frac{x}{58.44} (1/2)(278.1) + \frac{y}{174.27} (303.3) &= 21.75 \end{aligned}$$

We have two equations:

$$(2.379)x + (1.740)y = 21.75 \text{ and } x + y = 10.00. \text{ Solving:}$$

$$x = 6.81 \text{ g NaCl}; \quad \frac{6.81 \text{ g NaCl}}{10.00 \text{ g mixture}} \times 100 = 68.1\% \text{ NaCl}$$

101. a. Flow rate = $5.00 \times 10^4 \text{ L/s} + 3.50 \times 10^3 \text{ L/s} = 5.35 \times 10^4 \text{ L/s}$

$$\text{b. } C_{\text{HCl}} = \frac{3.50 \times 10^3 (65.0)}{5.35 \times 10^4} = 4.25 \text{ ppm HCl}$$

c. 1 ppm = 1 mg/kg H_2O = 1 mg/L (assuming density = 1.00 g/mL)

$$8.00 \text{ h} \times \frac{60 \text{ min}}{\text{h}} \times \frac{60 \text{ s}}{\text{min}} \times \frac{1.80 \times 10^4 \text{ L}}{\text{s}} \times \frac{4.25 \text{ mg HCl}}{\text{L}} \times \frac{1 \text{ g}}{1000 \text{ mg}} = 2.20 \times 10^6 \text{ g HCl}$$

$$2.20 \times 10^6 \text{ g HCl} \times \frac{1 \text{ mol HCl}}{36.46 \text{ g HCl}} \times \frac{1 \text{ mol CaO}}{2 \text{ mol HCl}} \times \frac{56.08 \text{ g Ca}}{\text{mol CaO}} = 1.69 \times 10^6 \text{ g CaO}$$

- d. The concentration of Ca^{2+} going into the second plant was:

$$\frac{5.00 \times 10^4(10.2)}{5.35 \times 10^4} = 9.53 \text{ ppm}$$

The second plant used: $1.80 \times 10^4 \text{ L/s} \times (8.00 \times 60 \times 60) \text{ s} = 5.18 \times 10^8 \text{ L}$ of water.

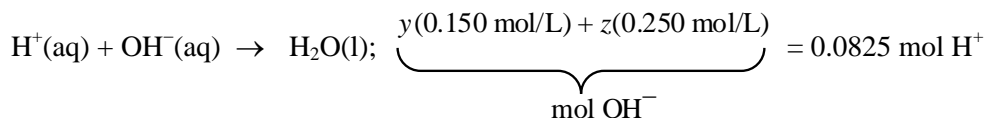
$$1.69 \times 10^6 \text{ g CaO} \times \frac{40.08 \text{ g Ca}^{2+}}{56.08 \text{ g CaO}} = 1.21 \times 10^6 \text{ g Ca}^{2+} \text{ was added to this water.}$$

$$C_{\text{Ca}^{2+}} (\text{plant water}) = 9.53 + \frac{1.21 \times 10^9 \text{ mg}}{5.18 \times 10^8 \text{ L}} = 9.53 + 2.34 = 11.87 \text{ ppm}$$

Because 90.0% of this water is returned, $(1.80 \times 10^4) \times 0.900 = 1.62 \times 10^4 \text{ L/s}$ of water with 11.87 ppm Ca^{2+} is mixed with $(5.35 - 1.80) \times 10^4 = 3.55 \times 10^4 \text{ L/s}$ of water containing 9.53 ppm Ca^{2+} .

$$C_{\text{Ca}^{2+}} (\text{final}) = \frac{(1.62 \times 10^4 \text{ L/s})(11.87 \text{ ppm}) + (3.55 \times 10^4 \text{ L/s})(9.53 \text{ ppm})}{1.62 \times 10^4 \text{ L/s} + 3.55 \times 10^4 \text{ L/s}} = 10.3 \text{ ppm}$$

102. $0.2750 \text{ L} \times 0.300 \text{ mol/L} = 0.0825 \text{ mol H}^+$; let $y = \text{volume (L) delivered by Y}$ and $z = \text{volume (L) delivered by Z}$.



$$0.2750 \text{ L} + y + z = 0.655 \text{ L}, \quad y + z = 0.380, \quad z = 0.380 - y$$

$$y(0.150) + (0.380 - y)(0.250) = 0.0825, \text{ solving: } y = 0.125 \text{ L}, \quad z = 0.255 \text{ L}$$

$$\text{Flow rate for Y} = \frac{125 \text{ mL}}{60.65 \text{ min}} = 2.06 \text{ mL/min}; \quad \text{flow rate for Z} = \frac{255 \text{ mL}}{60.65 \text{ min}} = 4.20 \text{ mL/min}$$

103. a. Let $x = \text{mass of Mg}$, so $10.00 - x = \text{mass of Zn}$. $\text{Ag}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \rightarrow \text{AgCl}(\text{s})$.

From the given balanced equations, there is a 2 : 1 mole ratio between mol Mg and mol Cl^- . The same is true for Zn. Because mol $\text{Ag}^+ = \text{mol Cl}^-$ present, one can setup an equation relating mol Cl^- present to mol Ag^+ added.

$$\begin{aligned} x \text{ g Mg} \times \frac{1 \text{ mol Mg}}{24.31 \text{ g Mg}} \times \frac{2 \text{ mol Cl}^-}{\text{mol Mg}} + (10.00 - x) \text{ g Zn} \times \frac{1 \text{ mol Zn}}{65.38 \text{ g Zn}} \times \frac{2 \text{ mol Cl}^-}{\text{mol Zn}} \\ = 0.156 \text{ L} \times \frac{3.00 \text{ mol Ag}^+}{\text{L}} \times \frac{1 \text{ mol Cl}^-}{\text{mol Ag}^+} = 0.468 \text{ mol Cl}^- \end{aligned}$$

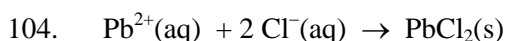
$$\frac{2x}{24.31} + \frac{2(10.00 - x)}{65.38} = 0.468, \quad 24.31 \times 65.38 \left(\frac{2x}{24.31} + \frac{20.00 - 2x}{65.38} = 0.468 \right)$$

$$(130.8)x + 486.2 - (48.62)x = 743.8 \text{ (carrying 1 extra significant figure)}$$

$$(82.2)x = 257.6, \quad x = 3.13 \text{ g Mg}; \quad \% \text{ Mg} = \frac{3.13 \text{ g Mg}}{10.00 \text{ g mixture}} \times 100 = 31.3\% \text{ Mg}$$

$$\text{b. } 0.156 \text{ L} \times \frac{3.00 \text{ mol Ag}^+}{\text{L}} \times \frac{1 \text{ mol Cl}^-}{\text{mol Ag}^+} = 0.468 \text{ mol Cl}^- = 0.468 \text{ mol HCl added}$$

$$M_{\text{HCl}} = \frac{0.468 \text{ mol}}{0.0780 \text{ L}} = 6.00 \text{ M HCl}$$



$$3.407 \text{ g PbCl}_2 \times \frac{1 \text{ mol PbCl}_2}{278.1 \text{ g PbCl}_2} \times \frac{1 \text{ mol Pb}^{2+}}{\text{mol PbCl}_2} = 0.01225 \text{ mol Pb}^{2+}$$

$$\frac{0.01225 \text{ mol}}{2.00 \times 10^{-3} \text{ L}} = 6.13 \text{ M Pb}^{2+} = 6.13 \text{ M Pb}(\text{NO}_3)_2$$

This is also the $\text{Pb}(\text{NO}_3)_2$ concentration in the 80.0 mL of evaporated solution.

$$\text{Original concentration} = \frac{\text{moles Pb}(\text{NO}_3)_2}{\text{original volume}} = \frac{0.0800 \text{ L} \times 6.13 \text{ mol/L}}{0.1000 \text{ L}} = 4.90 \text{ M Pb}(\text{NO}_3)_2$$



$$\text{Molar mass of PCB} = 12(12.01) + (10 - n)(1.008) + n(35.45) = 154.20 + (34.44)n$$

Because $n \text{ mol AgCl}$ are produced for every 1 mol PCB reacted, $n(143.4) \text{ g}$ of AgCl will be produced for every $[154.20 + (34.44)n] \text{ g}$ of PCB reacted.

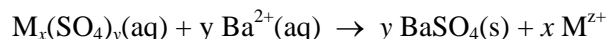
$$\frac{\text{Mass of AgCl}}{\text{Mass of PCB}} = \frac{(143.4)n}{154.20 + (34.44)n} \text{ or } \text{mass}_{\text{AgCl}}[154.20 + (34.44)n] = \text{mass}_{\text{PCB}}(143.4)n$$

$$\text{b. } 0.4971[154.20 + (34.44)n] = 0.1947(143.4)n, \quad 76.65 + (17.12)n = (27.92)n$$

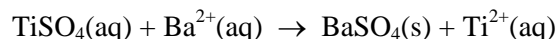
$$76.65 = (10.80)n, \quad n = 7.097$$

106. $\text{Mol BaSO}_4 = 0.2327 \text{ g} \times \frac{1 \text{ mol}}{233.4 \text{ g}} = 9.970 \times 10^{-4} \text{ mol BaSO}_4$

The moles of the sulfate salt depends on the formula of the salt. The general equation is:



Depending on the value of y , the mole ratio between the unknown sulfate salt and BaSO_4 varies. For example, if Pat thinks the formula is TiSO_4 , the equation becomes:

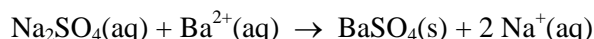


Because there is a 1 : 1 mole ratio between mol BaSO₄ and mol TiSO₄, you need 9.970×10^{-4} mol of TiSO₄. Because 0.1472 g of salt was used, the compound would have a molar mass of (assuming the TiSO₄ formula):

$$0.1472 \text{ g} / 9.970 \times 10^{-4} \text{ mol} = 147.6 \text{ g/mol}$$

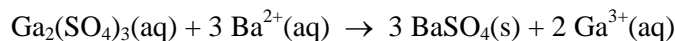
From atomic masses in the periodic table, the molar mass of TiSO₄ is 143.95 g/mol. From just these data, TiSO₄ seems reasonable.

Chris thinks the salt is sodium sulfate, which would have the formula Na₂SO₄. The equation is:



As with TiSO₄, there is a 1:1 mole ratio between mol BaSO₄ and mol Na₂SO₄. For sodium sulfate to be a reasonable choice, it must have a molar mass of about 147.6 g/mol. Using atomic masses, the molar mass of Na₂SO₄ is 142.05 g/mol. Thus Na₂SO₄ is also reasonable.

Randy, who chose gallium, deduces that gallium should have a 3+ charge (because it is in column 3A), and the formula of the sulfate would be Ga₂(SO₄)₃. The equation would be:



The calculated molar mass of Ga₂(SO₄)₃ would be:

$$\frac{0.1472 \text{ g Ga}_2(\text{SO}_4)_3}{9.970 \times 10^{-4} \text{ mol BaSO}_4} \times \frac{3 \text{ mol BaSO}_4}{\text{mol Ga}_2(\text{SO}_4)_3} = 442.9 \text{ g/mol}$$

Using atomic masses, the molar mass of Ga₂(SO₄)₃ is 427.65 g/mol. Thus Ga₂(SO₄)₃ is also reasonable.

Looking in references, sodium sulfate (Na₂SO₄) exists as a white solid with orthorhombic crystals, whereas gallium sulfate Ga₂(SO₄)₃ is a white powder. Titanium sulfate exists as a green powder, but its formula is Ti₂(SO₄)₃. Because this has the same formula as gallium sulfate, the calculated molar mass should be around 443 g/mol. However, the molar mass of Ti₂(SO₄)₃ is 383.97 g/mol. It is unlikely, then, that the salt is titanium sulfate.

To distinguish between Na₂SO₄ and Ga₂(SO₄)₃, one could dissolve the sulfate salt in water and add NaOH. Ga³⁺ would form a precipitate with the hydroxide, whereas Na₂SO₄ would not. References confirm that gallium hydroxide is insoluble in water.

107. There are three unknowns so we need three equations to solve for the unknowns. Let x = mass AgNO₃, y = mass CuCl₂, and z = mass FeCl₃. Then $x + y + z = 1.0000$ g. The Cl⁻ in CuCl₂ and FeCl₃ will react with the excess AgNO₃ to form the precipitate AgCl(s). Assuming silver has an atomic mass of 107.90:

$$\text{Mass of Cl in mixture} = 1.7809 \text{ g AgCl} \times \frac{35.45 \text{ g Cl}}{143.35 \text{ g AgCl}} = 0.4404 \text{ g Cl}$$

$$\text{Mass of Cl from CuCl}_2 = y \text{ g CuCl}_2 \times \frac{2(35.45) \text{ g Cl}}{134.45 \text{ g CuCl}_2} = (0.5273)y$$

$$\text{Mass of Cl from FeCl}_3 = z \text{ g FeCl}_3 \times \frac{3(35.45) \text{ g Cl}}{162.20 \text{ g FeCl}_3} = (0.6557)z$$

The second equation is: $0.4404 \text{ g Cl} = (0.5273)y + (0.6557)z$

Similarly, let's calculate the mass of metals in each salt.

$$\text{Mass of Ag in AgNO}_3 = x \text{ g AgNO}_3 \times \frac{107.9 \text{ g Ag}}{169.91 \text{ g AgNO}_3} = (0.6350)x$$

For CuCl_2 and FeCl_3 , we already calculated the amount of Cl in each initial amount of salt; the remainder must be the mass of metal in each salt.

$$\text{Mass of Cu in CuCl}_2 = y - (0.5273)y = (0.4727)y$$

$$\text{Mass of Fe in FeCl}_3 = z - (0.6557)z = (0.3443)z$$

The third equation is: $0.4684 \text{ g metals} = (0.6350)x + (0.4727)y + (0.3443)z$

We now have three equations with three unknowns. Solving:

$$\begin{array}{r} -0.6350(1.0000 = x + y + z) \\ 0.4684 = (0.6350)x + (0.4727)y + (0.3443)z \end{array}$$

$$-0.1666 = -(0.1623)y - (0.2907)z$$

$$\frac{0.5273}{0.1623} [-0.1666 = -(0.1623)y - (0.2907)z]$$

$$0.4404 = (0.5273)y + (0.6557)z$$

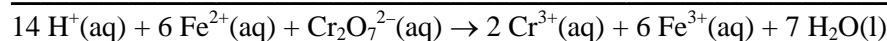
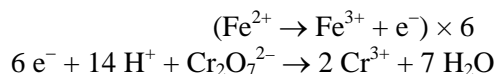
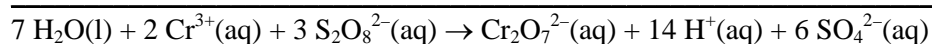
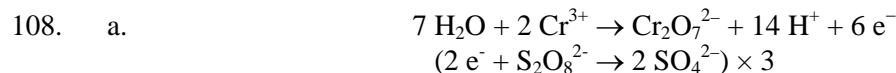
$$-0.1009 = -(0.2888)z, \quad z = \frac{0.1009}{0.2888} = 0.3494 \text{ g FeCl}_3$$

$$0.4404 = (0.5273)y + 0.6557(0.3494), \quad y = 0.4007 \text{ g CuCl}_2$$

$$x = 1.0000 - y - z = 1.0000 - 0.4007 - 0.3494 = 0.2499 \text{ g AgNO}_3$$

$$\text{Mass \% AgNO}_3 = \frac{0.2499 \text{ g}}{1.0000 \text{ g}} \times 100 = 24.99\% \text{ AgNO}_3$$

$$\text{Mass \% CuCl}_2 = \frac{0.4007 \text{ g}}{1.0000 \text{ g}} \times 100 = 40.07\% \text{ CuCl}_2; \quad \text{mass \% FeCl}_3 = 34.94\%$$



b.
$$8.58 \times 10^{-3} \text{ L} \times \frac{0.0520 \text{ mol Cr}_2\text{O}_7^{2-}}{\text{L}} \times \frac{6 \text{ mol Fe}^{2+}}{\text{mol Cr}_2\text{O}_7^{2-}} = 2.68 \times 10^{-3} \text{ mol of excess Fe}^{2+}$$

$$\text{Fe}^{2+} (\text{total}) = 3.000 \text{ g Fe}(\text{NH}_4)_2(\text{SO}_4)_2 \cdot 6\text{H}_2\text{O} \times \frac{1 \text{ mol}}{392.17 \text{ g}} = 7.650 \times 10^{-3} \text{ mol Fe}^{2+}$$

$7.650 \times 10^{-3} - 2.68 \times 10^{-3} = 4.97 \times 10^{-3} \text{ mol Fe}^{2+}$ reacted with $\text{Cr}_2\text{O}_7^{2-}$ generated from the Cr plating.

The Cr plating contained:

$$4.97 \times 10^{-3} \text{ mol Fe}^{2+} \times \frac{1 \text{ mol Cr}_2\text{O}_7^{2-}}{6 \text{ mol Fe}^{2+}} \times \frac{2 \text{ mol Cr}^{3+}}{\text{mol Cr}_2\text{O}_7^{2-}} = 1.66 \times 10^{-3} \text{ mol Cr}^{3+}$$

$$= 1.66 \times 10^{-3} \text{ mol Cr}$$

$$1.66 \times 10^{-3} \text{ mol Cr} \times \frac{52.00 \text{ g Cr}}{\text{mol Cr}} = 8.63 \times 10^{-2} \text{ g Cr}$$

$$\text{Volume of Cr plating} = 8.63 \times 10^{-2} \text{ g} \times \frac{1 \text{ cm}^3}{7.19 \text{ g}} = 1.20 \times 10^{-2} \text{ cm}^3 = \text{area} \times \text{thickness}$$

$$\text{Thickness of Cr plating} = \frac{1.20 \times 10^{-2} \text{ cm}^3}{40.0 \text{ cm}^2} = 3.00 \times 10^{-4} \text{ cm} = 300. \mu\text{m}$$



$$+3 + 2(+2) + 3x + 6.5(-2) = 0$$

$$7 + 3x - 13 = 0, \quad 3x = 6, \quad x = +2 \quad \text{Only Cu}^{2+} \text{ present.}$$

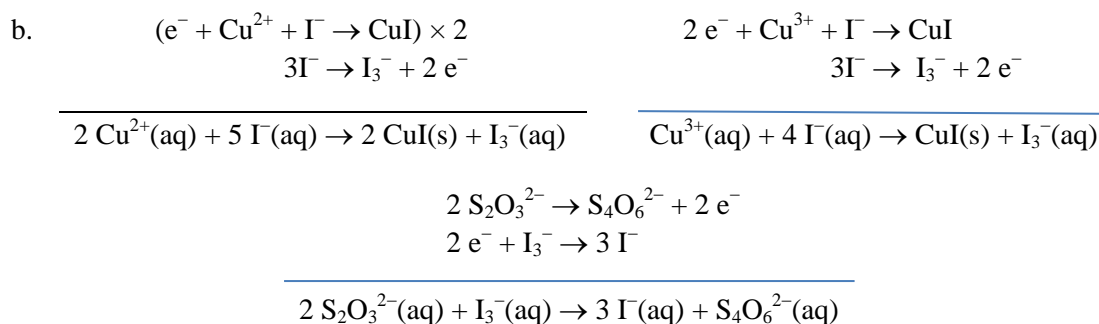


$$+3 + 2(+2) + 3x + 7(-2) = 0, \quad x = +2 \frac{1}{3} \text{ or } 2.33$$

This corresponds to two Cu^{2+} and one Cu^{3+} present.



$$+3 + 2(+2) + 3x + 8(-2) = 0, \quad x = +3; \quad \text{Only Cu}^{3+} \text{ present.}$$



c. Step II data: All Cu is converted to Cu^{2+} . *Note:* Superconductor abbreviated as "123".

$$22.57 \times 10^{-3} \text{ L} \times \frac{0.1000 \text{ mol S}_2\text{O}_3^{2-}}{\text{L}} \times \frac{1 \text{ mol I}_3^-}{2 \text{ mol S}_2\text{O}_3^{2-}} \times \frac{2 \text{ mol Cu}^{2+}}{\text{mol I}_3^-} = 2.257 \times 10^{-3} \text{ mol Cu}^{2+}$$

$$2.257 \times 10^{-3} \text{ mol Cu} \times \frac{1 \text{ mol "123"}}{3 \text{ mol Cu}} = 7.523 \times 10^{-4} \text{ mol "123"}$$

$$\text{Molar mass of YBa}_2\text{Cu}_3\text{O}_x = \frac{0.5402 \text{ g}}{7.523 \times 10^{-4} \text{ mol}} = 670.2 \text{ g/mol}$$

$$670.2 = 88.91 + 2(137.3) + 3(63.55) + x(16.00), \quad 670.2 = 554.2 + x(16.00)$$

$x = 7.250$; formula is $\text{YBa}_2\text{Cu}_3\text{O}_{7.25}$.

Check with Step I data: Both Cu^{2+} and Cu^{3+} present.

$$37.77 \times 10^{-3} \text{ L} \times \frac{0.1000 \text{ mol S}_2\text{O}_3^{2-}}{\text{L}} \times \frac{1 \text{ mol I}_3^-}{2 \text{ mol S}_2\text{O}_3^{2-}} = 1.889 \times 10^{-3} \text{ mol I}_3^-$$

We get 1 mol I_3^- per mol Cu^{3+} and 1 mol I_3^- per 2 mol Cu^{2+} . Let $n_{\text{Cu}^{3+}} = \text{mol Cu}^{3+}$ and $n_{\text{Cu}^{2+}} = \text{mol Cu}^{2+}$, then:

$$n_{\text{Cu}^{3+}} + \frac{n_{\text{Cu}^{2+}}}{2} = 1.889 \times 10^{-3} \text{ mol}$$

In addition: $\frac{0.5625 \text{ g}}{670.2 \text{ g/mol}} = 8.393 \times 10^{-4} \text{ mol "123"}$; this amount of "123" contains:

$$3(8.393 \times 10^{-4}) = 2.518 \times 10^{-3} \text{ mol Cu total} = n_{\text{Cu}^{3+}} + n_{\text{Cu}^{2+}}$$

Solving by simultaneous equations:

$$\begin{array}{r}
 n_{\text{Cu}^{3+}} + n_{\text{Cu}^{2+}} = 2.518 \times 10^{-3} \\
 -n_{\text{Cu}^{3+}} - \frac{n_{\text{Cu}^{2+}}}{2} = -1.889 \times 10^{-3} \\
 \hline
 \end{array}$$

$$\frac{n_{\text{Cu}^{2+}}}{2} = 6.29 \times 10^{-4}$$

$$n_{\text{Cu}^{2+}} = 1.26 \times 10^{-3} \text{ mol Cu}^{2+}; \quad n_{\text{Cu}^{3+}} = 2.518 \times 10^{-3} - 1.26 \times 10^{-3} = 1.26 \times 10^{-3} \text{ mol Cu}^{3+}$$

This sample of superconductor contains equal moles of Cu^{2+} and Cu^{3+} . Therefore, 1 mole of $\text{YBa}_2\text{Cu}_3\text{O}_x$ contains 1.50 mol Cu^{2+} and 1.50 mol Cu^{3+} . Solving for x using oxidation states:

$$+3 + 2(+2) + 1.50(+2) + 1.50(+3) + x(-2) = 0, \quad 14.50 = 2x, \quad x = 7.25$$

The two experiments give the same result, $x = 7.25$ with formula $\text{YBa}_2\text{Cu}_3\text{O}_{7.25}$.

Average oxidation state of Cu:

$$+3 + 2(+2) + 3(x) + 7.25(-2) = 0, \quad 3x = 7.50, \quad x = +2.50$$

As determined from Step I data, this superconductor sample contains equal moles of Cu^{2+} and Cu^{3+} , giving an average oxidation state of +2.50.

$$110. \quad 0.298 \text{ g BaSO}_4 \times \frac{96.07 \text{ g SO}_4^{2-}}{233.4 \text{ g BaSO}_4} = 0.123 \text{ g SO}_4^{2-}; \quad \% \text{ sulfate} = \frac{0.123 \text{ g SO}_4^{2-}}{0.205 \text{ g}} = 60.0\%$$

Assume we have 100.0 g of the mixture of Na_2SO_4 and K_2SO_4 . There are:

$$60.0 \text{ g SO}_4^{2-} \times \frac{1 \text{ mol}}{96.07 \text{ g}} = 0.625 \text{ mol SO}_4^{2-}$$

There must be $2 \times 0.625 = 1.25$ mol of 1+ cations to balance the 2- charge of SO_4^{2-} .

Let x = number of moles of K^+ and y = number of moles of Na^+ ; then $x + y = 1.25$.

The total mass of Na^+ and K^+ must be 40.0 g in the assumed 100.0 g of mixture. Setting up an equation:

$$x \text{ mol K}^+ \times \frac{39.10 \text{ g}}{\text{mol}} + y \text{ mol Na}^+ \times \frac{22.99 \text{ g}}{\text{mol}} = 40.0 \text{ g}$$

So, we have two equations with two unknowns: $x + y = 1.25$ and $(39.10)x + (22.99)y = 40.0$

$$x = 1.25 - y, \quad \text{so } 39.10(1.25 - y) + (22.99)y = 40.0$$

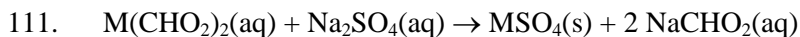
$$48.9 - (39.10)y + (22.99)y = 40.0, \quad -(16.11)y = -8.9$$

$$y = 0.55 \text{ mol Na}^+ \text{ and } x = 1.25 - 0.55 = 0.70 \text{ mol K}^+$$

Therefore:

$$0.70 \text{ mol K}^+ \times \frac{1 \text{ mol K}_2\text{SO}_4}{2 \text{ mol K}^+} = 0.35 \text{ mol K}_2\text{SO}_4; \quad 0.35 \text{ mol K}_2\text{SO}_4 \times \frac{174.27 \text{ g}}{\text{mol}} = 61 \text{ g K}_2\text{SO}_4$$

We assumed 100.0 g; therefore, the mixture is 61% K_2SO_4 and 39% Na_2SO_4 .

Marathon Problems

From the balanced molecular equation, the moles of $M(\text{CHO}_2)_2$ present initially must equal the moles of $\text{MSO}_4(\text{s})$ formed. Because moles = mass/molar mass and letting A_M = the atomic mass of M:

$$\text{mol MSO}_4 = \frac{\text{mass MSO}_4}{\text{molar mass of MSO}_4} = \frac{9.9392 \text{ g}}{A_M + 96.07}$$

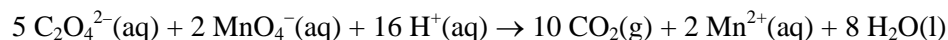
$$\text{mol M}(\text{CHO}_2)_2 = \frac{\text{mass M}(\text{CHO}_2)_2}{\text{molar mass of M}(\text{CHO}_2)_2} = \frac{9.7416 \text{ g}}{A_M + 90.04}$$

Because $\text{mol MSO}_4 = \text{mol M}(\text{CHO}_2)_2$:

$$\frac{9.9392}{A_M + 96.07} = \frac{9.7416}{A_M + 90.04}, \quad (9.9392)A_M + 894.9 = (9.7416)A_M + 935.9$$

$$A_M = \frac{41.0}{0.1976} = 207; \quad \text{from the periodic table, the unknown element M is Pb.}$$

From the information in the second paragraph, we can determine the concentration of the KMnO_4 solution. Using the half-reaction method, the balanced reaction between MnO_4^- and $\text{C}_2\text{O}_4^{2-}$ is:



$$\begin{aligned} \text{Mol MnO}_4^- &= 0.9234 \text{ g Na}_2\text{C}_2\text{O}_4 \times \frac{1 \text{ mol Na}_2\text{C}_2\text{O}_4}{134.00 \text{ g}} \times \frac{1 \text{ mol C}_2\text{O}_4^{2-}}{1 \text{ mol Na}_2\text{C}_2\text{O}_4} \times \frac{2 \text{ mol MnO}_4^-}{5 \text{ mol C}_2\text{O}_4^{2-}} \\ &= 2.756 \times 10^{-3} \text{ mol MnO}_4^- \end{aligned}$$

$$M_{\text{KMnO}_4} = M_{\text{MnO}_4^-} = \frac{\text{mol MnO}_4^-}{\text{volume}} = \frac{2.756 \times 10^{-3} \text{ mol}}{0.01855 \text{ L}} = 0.1486 \text{ mol/L}$$

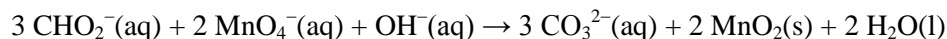
From the third paragraph, the standard KMnO_4 solution reacts with formate ions from the filtrate. We must determine the moles of CHO_2^- ions present in order to determine volume of KMnO_4 solution. The moles of CHO_2^- ions present initially are:

$$9.7416 \text{ g Pb}(\text{CHO}_2)_2 \times \frac{\text{mol Pb}(\text{CHO}_2)_2}{297.2 \text{ g}} \times \frac{2 \text{ mol CHO}_2^-}{1 \text{ mol Pb}(\text{CHO}_2)_2} = 6.556 \times 10^{-2} \text{ mol CHO}_2^-$$

The moles of CHO_2^- present in 10.00 mL of diluted solution are:

$$0.01000 \text{ L} \times \frac{6.556 \times 10^{-2} \text{ mol CHO}_2^-}{0.2500 \text{ L}} = 2.622 \times 10^{-3} \text{ mol CHO}_2^-$$

Using the half-reaction method in basic solution, the balanced reaction between CHO_2^- and MnO_4^- is:



Determining the volume of MnO_4^- solution:

$$2.622 \times 10^{-3} \text{ mol CHO}_2^- \times \frac{2 \text{ mol MnO}_4^-}{3 \text{ mol CHO}_2^-} \times \frac{1 \text{ L}}{0.1486 \text{ mol MnO}_4^-} = 1.176 \times 10^{-2} \text{ L} \\ = 11.76 \text{ mL}$$

The titration requires 11.76 mL of the standard KMnO_4 solution.

112. a. Compound A = $\text{M}(\text{NO}_3)_x$; in 100.00 g of compd.: $8.246 \text{ g N} \times \frac{48.00 \text{ g O}}{14.01 \text{ g N}} = 28.25 \text{ g O}$

Thus the mass of nitrate in the compound = $8.246 + 28.25 \text{ g} = 36.50 \text{ g}$ (if $x = 1$).

If $x = 1$: mass of M = $100.00 - 36.50 \text{ g} = 63.50 \text{ g}$

$$\text{Mol M} = \text{mol N} = \frac{8.246 \text{ g}}{14.01 \text{ g/mol}} = 0.5886 \text{ mol}$$

$$\text{Molar mass of metal M} = \frac{63.50 \text{ g}}{0.5886 \text{ mol}} = 107.9 \text{ g/mol (This is silver, Ag.)}$$

If $x = 2$: mass of M = $100.00 - 2(36.50) = 27.00 \text{ g}$

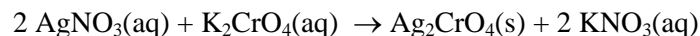
$$\text{Mol M} = \frac{1}{2} \text{ mol N} = \frac{0.5886 \text{ mol}}{2} = 0.2943 \text{ mol}$$

$$\text{Molar mass of metal M} = \frac{27.00 \text{ g}}{0.2943 \text{ mol}} = 91.74 \text{ g/mol}$$

This is close to Zr, but Zr does not form stable 2+ ions in solution; it forms stable 4+ ions. Because we cannot have $x = 3$ or more nitrates (three nitrates would have a mass greater than 100.00 g), compound A must be AgNO_3 .

Compound B: K_2CrO_x is the formula. This salt is composed of K^+ and CrO_x^{2-} ions. Using oxidation states, $6 + x(-2) = -2$, $x = 4$. Compound B is K_2CrO_4 (potassium chromate).

b. The reaction is:



The blood red precipitate is $\text{Ag}_2\text{CrO}_4(\text{s})$.

- c. 331.8 g Ag_2CrO_4 formed; this is equal to the molar mass of Ag_2CrO_4 , so 1 mole of precipitate formed. From the balanced reaction, we need 2 mol AgNO_3 to react with 1 mol K_2CrO_4 to produce 1 mol (331.8 g) of Ag_2CrO_4 .

$$2.000 \text{ mol AgNO}_3 \times \frac{169.9 \text{ g}}{\text{mol}} = 339.8 \text{ g AgNO}_3$$

$$1.000 \text{ mol K}_2\text{CrO}_4 \times \frac{194.2 \text{ g}}{\text{mol}} = 194.2 \text{ g K}_2\text{CrO}_4$$

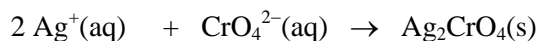
The problem says that we have equal masses of reactants. Our two choices are 339.8 g $\text{AgNO}_3 + 339.8 \text{ g K}_2\text{CrO}_4$ or 194.2 g $\text{AgNO}_3 + 194.2 \text{ g K}_2\text{CrO}_4$. If we assume the 194.2 g quantities are correct, then when 194.2 g K_2CrO_4 (1 mol) reacts, 339.8 g AgNO_3 (2.0 mol) must be present to react with all the K_2CrO_4 . We only have 194.2 g AgNO_3 present; this cannot be correct. Instead of K_2CrO_4 limiting, AgNO_3 must be limiting, and we have reacted 339.8 g AgNO_3 and 339.8 g K_2CrO_4 .

$$\text{Solution A: } \frac{2.000 \text{ mol Ag}^+}{0.5000 \text{ L}} = 4.000 \text{ M Ag}^+; \quad \frac{2.000 \text{ mol NO}_3^-}{0.5000 \text{ L}} = 4.000 \text{ M NO}_3^-$$

$$\text{Solution B: } 339.8 \text{ g K}_2\text{CrO}_4 \times \frac{1 \text{ mol}}{194.2 \text{ g}} = 1.750 \text{ mol K}_2\text{CrO}_4$$

$$\frac{2 \times 1.750 \text{ mol K}^+}{0.5000 \text{ L}} = 7.000 \text{ M K}^+; \quad \frac{1.750 \text{ mol CrO}_4^{2-}}{0.5000 \text{ L}} = 3.500 \text{ M CrO}_4^{2-}$$

- d. After the reaction, moles of K^+ and moles of NO_3^- remain unchanged because they are spectator ions. Because Ag^+ is limiting, its concentration will be 0 M after precipitation is complete.



Initial	2.000 mol	1.750 mol	0
Change	-2.000 mol	-1.000 mol	+1.000 mol
After rxn	0	0.750 mol	1.000 mol

$$M_{\text{K}^+} = \frac{2 \times 1.750 \text{ mol}}{1.0000 \text{ L}} = 3.500 \text{ M K}^+; \quad M_{\text{NO}_3^-} = \frac{2.000 \text{ mol}}{1.0000 \text{ L}} = 2.000 \text{ M NO}_3^-$$

$$M_{\text{CrO}_4^{2-}} = \frac{0.750 \text{ mol}}{1.0000 \text{ L}} = 0.750 \text{ M CrO}_4^{2-}; \quad M_{\text{Ag}^+} = 0 \text{ M (the limiting reagent)}$$