

- 4.66 *Analyze.* Given: BAC (definition from Exercise 4.65), vol of blood. Find: mass alcohol in bloodstream.

Plan. Change BAC (g/100 mL) to (g/L), then times vol of blood in L.

Solve. BAC = 0.10 g/100 mL

$$\frac{0.10 \text{ g alcohol}}{100 \text{ mL blood}} \times \frac{1000 \text{ mL}}{1 \text{ L}} \times 5.0 \text{ L blood} = 5.0 \text{ g alcohol}$$

- 4.67 *Plan.* Proceed as in Sample Exercises 4.13.

$$M = \frac{\text{mol}}{\text{L}}; \text{ mol} = \frac{\text{g}}{\text{MM}} \quad (\text{MM is the symbol for molar mass in this manual.})$$

Solve.

$$(a) \quad \frac{6.86 \text{ mol CH}_3\text{CH}_2\text{OH}}{1 \text{ L}} \times 1.00 \text{ L} \times \frac{46.07 \text{ g CH}_3\text{CH}_2\text{OH}}{1 \text{ mol CH}_3\text{CH}_2\text{OH}} = 316.04 = 316 \text{ g CH}_3\text{CH}_2\text{OH}$$

Check. $(7 \times 50) \approx 350$ g ethanol (this is an upper limit)

$$(b) \quad 316.04 \text{ g CH}_3\text{CH}_2\text{OH} \times \frac{1 \text{ mL}}{0.789 \text{ g CH}_3\text{CH}_2\text{OH}} = 400.56 = 401 \text{ mL} = 0.401 \text{ L CH}_3\text{CH}_2\text{OH}$$

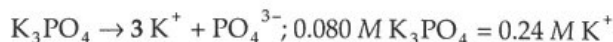
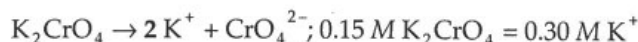
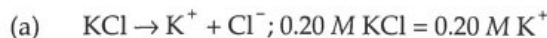
Check. $(320/0.8) \approx 400$ mL ethanol

- 4.68 $M = \frac{\text{mol}}{\text{L}}; \text{ mol} = \frac{\text{g}}{\text{MM}}$ (MM is the symbol for molar mass in this manual.)

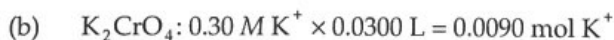
$$M = \frac{\text{mol solute}}{\text{L solution}}; \frac{124 \text{ mg C}_6\text{H}_8\text{O}_6}{0.2366 \text{ L}} \times \frac{1 \text{ g}}{1000 \text{ mg}} \times \frac{1 \text{ mol C}_6\text{H}_8\text{O}_6}{176.12 \text{ g C}_6\text{H}_8\text{O}_6} = 2.98 \times 10^{-3} \text{ M C}_6\text{H}_8\text{O}_6$$

- 4.69 *Analyze.* Given: formula and concentration of each solute. Find: concentration of K^+ in each solution. *Plan.* Note mol K^+ /mol solute and compare concentrations or total moles.

Solve.



0.15 M K_2CrO_4 has the highest K^+ concentration.



30.0 mL of 0.15 M K_2CrO_4 has more K^+ ions.

- 4.70 (a) 0.10 M $\text{BaI}_2 = 0.2 \text{ M I}^-$; 0.25 M $\text{KI} = 0.25 \text{ M I}^-$

0.25 M KI has the higher I^- concentration.

- (b) 0.10 M $\text{KI} = 0.1 \text{ M I}^-$; 0.040 M $\text{ZnI}_2 = 0.080 \text{ M I}^-$; 0.10 M KI has a higher I^- concentration than 0.040 M ZnI_2 . Total volume does not affect concentration.

$$(c) \quad 3.2 \text{ M HI} = 3.2 \text{ M I}^-$$

$$145 \text{ g NaI} \times \frac{1 \text{ mol NaI}}{149.9 \text{ g NaI}} \times \frac{1}{0.150 \text{ L}} = 6.45 \text{ M NaI} = 6.45 \text{ M I}^-$$

The NaI solution has the higher I^- concentration.

4.71 *Analyze.* Given: molecular formula and solution molarity. Find: concentration (M) of each ion.

Plan. Follow the logic in Sample Exercise 4.12.

Solve.



(d) *Plan.* There is no reaction between NaCl and $(\text{NH}_4)_2\text{CO}_3$, so this is just a dilution problem, $M_1V_1 = M_2V_2$. Then account for ion stoichiometry.

Solve. $45.0 \text{ mL} + 65.0 \text{ mL} = 110.0 \text{ mL}$ total volume

$$\frac{0.272 \text{ M NaCl} \times 45.0 \text{ mL}}{110.0 \text{ mL}} = 0.111 \text{ M NaCl}; 0.111 \text{ M Na}^+, 0.111 \text{ M Cl}^-$$

$$\frac{0.0247 \text{ M } (\text{NH}_4)_2\text{CO}_3 \times 65.0 \text{ mL}}{110.0 \text{ mL}} = 0.0146 \text{ M } (\text{NH}_4)_2\text{CO}_3$$

$$2 \times (0.0146 \text{ M}) = 0.0292 \text{ M NH}_4^+, 0.0146 \text{ M CO}_3^{2-}$$

Check. By adding the two solutions (with no common ions or chemical reaction), we have approximately doubled the solution volume, and reduced the concentration of each ion by approximately a factor of two.

4.72 (a) *Plan.* These two solutions have common ions. Find the ion concentration resulting from each solution, then add.

Solve. total volume = $42.0 \text{ mL} + 37.6 \text{ mL} = 79.6 \text{ mL}$

$$\frac{0.170 \text{ M NaOH} \times 42.0 \text{ mL}}{79.6 \text{ mL}} = 0.08970 = 0.0897 \text{ M NaOH};$$

$$0.0897 \text{ M Na}^+, 0.0897 \text{ M OH}^-$$

$$\frac{0.400 \text{ M NaOH} \times 37.6 \text{ mL}}{79.6 \text{ mL}} = 0.18894 = 0.189 \text{ M NaOH};$$

$$0.189 \text{ M Na}^+, 0.189 \text{ M OH}^-$$

$$\text{M Na}^+ = 0.08970 \text{ M} + 0.18894 \text{ M} = 0.27864 = 0.279 \text{ M Na}^+$$

$$\text{M OH}^- = \text{M Na}^+ = 0.279 \text{ M OH}^-$$

(b) *Plan.* No common ions; just dilution.

Solve. $44.0 \text{ mL} + 25.0 \text{ mL} = 69.0 \text{ mL}$

$$\frac{0.100 \text{ M Na}_2\text{SO}_4 \times 44.0 \text{ mL}}{69.0 \text{ mL}} = 0.06377 = 0.0638 \text{ M Na}_2\text{SO}_4$$

$$2 \times (0.06377 \text{ M}) = 0.1275 = 0.128 \text{ M Na}^+; 0.0638 \text{ M SO}_4^{2-}$$

$$\frac{0.150 \text{ M KCl} \times 25.0 \text{ mL}}{69.0 \text{ mL}} = 0.054348 = 0.0543 \text{ M KCl}$$

$$0.0543 \text{ M K}^+, 0.0543 \text{ M Cl}^-$$

(c) *Plan.* Calculate concentration of K^+ and Cl^- due to the added solid. Then sum to get total concentration of Cl^- .

$$\text{Solve. } \frac{3.60 \text{ g KCl}}{75.0 \text{ mL soln}} \times \frac{1 \text{ mol KCl}}{74.55 \text{ g KCl}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 0.6439 = 0.644 \text{ M KCl}$$

$$0.250 \text{ M CaCl}_2; 2(0.250 \text{ M}) = 0.500 \text{ M Cl}^-$$

$$\text{total Cl}^- = 0.644 \text{ M} + 0.500 \text{ M} = 1.144 \text{ M Cl}^-, 0.644 \text{ M K}^+, 0.250 \text{ M Ca}^{2+}$$

4.73 *Analyze/Plan.* Follow the logic of Sample Exercise 4.14.

Solve.

$$(a) \quad V_1 = M_2 V_2 / M_1; \frac{0.250 \text{ M NH}_3 \times 1000.0 \text{ mL}}{14.8 \text{ M NH}_3} = 16.89 = 16.9 \text{ mL } 14.8 \text{ M NH}_3$$

$$\text{Check. } 250/15 \approx 15 \text{ M}$$

$$(b) \quad M_2 = M_1 V_1 / V_2; \frac{14.8 \text{ M NH}_3 \times 10.0 \text{ mL}}{500 \text{ mL}} = 0.296 \text{ M NH}_3$$

$$\text{Check. } 150/500 \approx 0.30 \text{ M}$$

$$4.74 \quad (a) \quad V_1 = M_2 V_2 / M_1; \frac{0.500 \text{ M HNO}_3 \times 0.110 \text{ L}}{6.0 \text{ M HNO}_3} = 0.00917 \text{ L} = 9.2 \text{ mL } 6.0 \text{ M HNO}_3$$

$$(b) \quad M_2 = M_1 V_1 / V_2; \frac{6.0 \text{ M HNO}_3 \times 10.0 \text{ mL}}{250 \text{ mL}} = 0.240 \text{ M HNO}_3$$

4.75 *Analyze/Plan.* Calculate the number of drug molecules in 1.00 mL of the stock solution, using $M \times L = \text{moles}$ and Avogadro's number. Then calculate the desired ratio.

Solve. $1.00 \text{ mL} = 0.00100 \text{ L}$

$$\frac{1.5 \times 10^{-9} \text{ mol}}{\text{L}} \times 0.0010 \text{ L} \times \frac{6.022 \times 10^{23} \text{ molecules}}{\text{mole}} = 9.033 \times 10^{11} = 9.0 \times 10^{11} \text{ molecules}$$

$$\frac{9.033 \times 10^{11} \text{ drug molecules}}{2.0 \times 10^5 \text{ cancer cells}} = 4.517 \times 10^6 = 4.5 \times 10^6$$

- 4.76 *Analyze/Plan.* The 25.00 mL of antibiotic solution needs to contain a minimum of 1.0×10^8 molecules of the drug. Calculate the moles of drug this represents. The concentration of the stock solution is $5.00 \times 10^{-9} M$. Then,
 L stock solution = mol drug / M solution; L = mol / $5.00 \times 10^{-9} M$.

$$1.0 \times 10^8 \text{ molecules} \times \frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ molecules}} \times \frac{1 \text{ L}}{5.00 \times 10^{-9} \text{ mol}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 3.3 \times 10^{-5} \text{ mL}$$

A volume of 3.3×10^{-5} mL corresponds to 33×10^{-9} L, or 33 nL (nanoliters).

Two points are of note. First, desired results can be achieved with a very small amount of the drug, which reduces the cost. And, delivering such a small volume of stock solution may be a challenge. A dilution scheme for the stock solution could be employed. If 1.00 mL of the $5.00 \times 10^{-9} M$ antibiotic stock solution is diluted to 1.00 L, and 1.00 mL of this solution is diluted to 0.500 L, the resulting concentration of the diluted stock solution is then $1.00 \times 10^{-14} M$. Using the relationship $L = \text{mol} / M$, we find that 0.017 L or 17 mL of the $1.00 \times 10^{-14} M$ stock solution diluted to 25.00 mL would kill the desired amount of bacteria.

- 4.77 *Analyze.* Given: density of pure acetic acid, volume pure acetic acid, volume new solution. Find: molarity of new solution. *Plan.* Calculate the mass of acetic acid, CH_3COOH , present in 20.0 mL of the pure liquid. *Solve.*

$$20.00 \text{ mL acetic acid} \times \frac{1.049 \text{ g acetic acid}}{1 \text{ mL acetic acid}} = 20.98 \text{ g acetic acid}$$

$$20.98 \text{ g CH}_3\text{COOH} \times \frac{1 \text{ mol CH}_3\text{COOH}}{60.05 \text{ g CH}_3\text{COOH}} = 0.349375 = 0.3494 \text{ mol CH}_3\text{COOH}$$

$$M = \text{mol/L} = \frac{0.349375 \text{ mol CH}_3\text{COOH}}{0.2500 \text{ L solution}} = 1.39750 = 1.398 M \text{ CH}_3\text{COOH}$$

Check. $(20 \times 1) \approx 20$ g acid; $(20/60) \approx 0.33$ mol acid; $(0.33/0.25 = 0.33 \times 4) \approx 1.33 M$

- 4.78 $50.000 \text{ mL glycerol} \times \frac{1.2656 \text{ g glycerol}}{1 \text{ mL glycerol}} = 63.280 \text{ g glycerol}$

$$63.280 \text{ g C}_3\text{H}_8\text{O}_3 \times \frac{1 \text{ mol C}_3\text{H}_8\text{O}_3}{92.094 \text{ g C}_3\text{H}_8\text{O}_3} = 0.687124 = 0.68712 \text{ mol C}_3\text{H}_8\text{O}_3$$

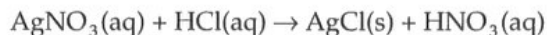
$$M = \frac{0.687124 \text{ mol C}_3\text{H}_8\text{O}_3}{0.25000 \text{ L solution}} = 2.7485 M \text{ C}_3\text{H}_8\text{O}_3$$

Solution Stoichiometry and Chemical Analysis (Section 4.6)

- 4.79 *Analyze.* Given: volume and molarity AgNO_3 , molarity HCl. Find: volume HCl or mass of KCl.

(a) *Plan.* $M \times L = \text{mol AgNO}_3 = \text{mol Ag}^+$; balanced equation gives ratio mol HCl/mol AgNO_3 ; mol HCl \rightarrow vol HCl. *Solve.*

$$\frac{0.200 \text{ mol AgNO}_3}{1 \text{ L}} \times 0.0150 \text{ L} = 3.00 \times 10^{-3} \text{ mol AgNO}_3(\text{aq})$$



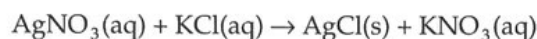
$$\text{mol HCl} = \text{mol AgNO}_3 = 3.00 \times 10^{-3} \text{ mol KCl}$$

$$3.00 \times 10^{-3} \text{ mol HCl} \times \frac{1 \text{ L}}{0.150 \text{ mol HCl}} = 0.0200 \text{ L} = 20.0 \text{ mL } 0.15 \text{ M HCl}$$

Check. $(0.2 \times 0.015) = 0.003 \text{ mol}$; $(0.003/0.15) \approx 0.02 \text{ L HCl}$

- (b) *Plan.* $M \times L = \text{mol AgNO}_3 = \text{mol Ag}^+$; balanced equation gives ratio mol KCl/mol AgNO₃; mol KCl → vol KCl. *Solve.*

$$\frac{0.200 \text{ mol AgNO}_3}{1 \text{ L}} \times 0.0150 \text{ L} = 3.00 \times 10^{-3} \text{ mol AgNO}_3(\text{aq})$$



$$\text{mol KCl} = \text{mol AgNO}_3 = 3.00 \times 10^{-3} \text{ mol KCl}$$

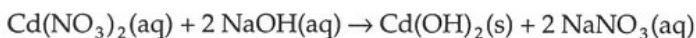
$$3.00 \times 10^{-3} \text{ mol KCl} \times \frac{74.55 \text{ g KCl}}{1 \text{ mol KCl}} = 0.224 \text{ g KCl}$$

Check. $(0.2 \times 0.015) = 0.003 \text{ mol}$; $(0.003 \times 75) \approx 0.225 \text{ g KCl}$

- (c) Clearly, the KCl reagent is virtually free relative to the HCl solution. The KCl analysis is more cost-effective.

- 4.80 *Plan.* $M \times L = \text{mol Cd}(\text{NO}_3)_2$; balanced equation → mol ratio → mol NaOH → g NaOH

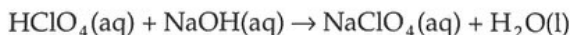
$$\text{Solve. } \frac{0.500 \text{ mol Cd}(\text{NO}_3)_2}{1 \text{ L}} \times 0.0350 \text{ L} = 0.0175 \text{ mol Cd}(\text{NO}_3)_2$$



$$0.0175 \text{ mol Cd}(\text{NO}_3)_2 \times \frac{2 \text{ mol NaOH}}{1 \text{ mol Cd}(\text{NO}_3)_2} \times \frac{40.00 \text{ g NaOH}}{1 \text{ mol NaOH}} = 1.40 \text{ g NaOH}$$

- 4.81 (a) *Analyze.* Given: M and vol base, M acid. Find: vol acid

Plan/Solve. Write the balanced equation for the reaction in question:



Calculate the moles of the known substance, in this case NaOH.

$$\text{moles NaOH} = M \times L = \frac{0.0875 \text{ mol NaOH}}{1 \text{ L}} \times 0.0500 \text{ L} = 0.004375 = 0.00438 \text{ mol NaOH}$$

Apply the mole ratio (mol unknown/mol known) from the chemical equation.

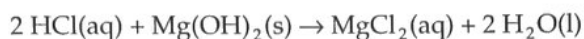
$$0.004375 \text{ mol NaOH} \times \frac{1 \text{ mol HClO}_4}{1 \text{ mol NaOH}} = 0.004375 \text{ mol HClO}_4$$

Calculate the desired quantity of unknown, in this case the volume of 0.115 M HClO₄ solution.

$$L = \text{mol}/M; L = 0.004375 \text{ mol HClO}_4 \times \frac{1 \text{ L}}{0.115 \text{ mol HClO}_4} = 0.0380 \text{ L} = 38.0 \text{ mL}$$

Check. $(0.09 \times 0.05) = 0.0045 \text{ mol}$; $(0.0045/0.11) \approx 0.040 \text{ L} \approx 40 \text{ mL}$

(b) Following the logic outlined in part (a):



$$2.87 \text{ g Mg(OH)}_2 \times \frac{1 \text{ mol Mg(OH)}_2}{58.32 \text{ g Mg(OH)}_2} = 0.049211 = 0.0492 \text{ mol Mg(OH)}_2$$

$$0.0492 \text{ mol Mg(OH)}_2 \times \frac{2 \text{ mol HCl}}{1 \text{ mol Mg(OH)}_2} = 0.0984 \text{ mol HCl}$$

$$L = \text{mol}/M = 0.09840 \text{ mol HCl} \times \frac{1 \text{ L HCl}}{0.128 \text{ mol HCl}} = 0.769 \text{ L} = 769 \text{ mL}$$

(c) $\text{AgNO}_3\text{(aq)} + \text{KCl(aq)} \rightarrow \text{AgCl(s)} + \text{KNO}_3\text{(aq)}$

$$785 \text{ mg KCl} \times \frac{1 \times 10^{-3} \text{ g}}{1 \text{ mg}} \times \frac{1 \text{ mol KCl}}{74.55 \text{ g KCl}} \times \frac{1 \text{ mol AgNO}_3}{1 \text{ mol KCl}} = 0.01053 = 0.0105 \text{ mol AgNO}_3$$

$$M = \text{mol}/L = \frac{0.01053 \text{ mol AgNO}_3}{0.0258 \text{ L}} = 0.408 \text{ M AgNO}_3$$

(d) $\text{HCl(aq)} + \text{KOH(aq)} \rightarrow \text{KCl(aq)} + \text{H}_2\text{O(l)}$

$$\frac{0.108 \text{ mol HCl}}{1 \text{ L}} \times 0.0453 \text{ L} \times \frac{1 \text{ mol KOH}}{1 \text{ mol HCl}} \times \frac{56.11 \text{ g KOH}}{1 \text{ mol KOH}} = 0.275 \text{ g KOH}$$

4.82 (a) $2 \text{HCl(aq)} + \text{Ba(OH)}_2\text{(aq)} \rightarrow \text{BaCl}_2\text{(aq)} + 2 \text{H}_2\text{O(l)}$

$$\frac{0.101 \text{ mol Ba(OH)}_2}{1 \text{ L Ba(OH)}_2} \times 0.0500 \text{ L Ba(OH)}_2 \times \frac{2 \text{ mol HCl}}{1 \text{ mol Ba(OH)}_2} \times \frac{1 \text{ L HCl}}{0.120 \text{ mol HCl}} = 0.0842 \text{ L or } 84.2 \text{ mL HCl soln}$$

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

$$0.200 \text{ g NaOH} \times \frac{1 \text{ mol NaOH}}{40.00 \text{ g NaOH}} \times \frac{1 \text{ mol H}_2\text{SO}_4}{2 \text{ mol NaOH}} \times \frac{1 \text{ L H}_2\text{SO}_4}{0.125 \text{ mol H}_2\text{SO}_4} = 0.0200 \text{ L or } 20.0 \text{ mL H}_2\text{SO}_4 \text{ soln}$$

(c) $\text{BaCl}_2\text{(aq)} + \text{Na}_2\text{SO}_4\text{(aq)} \rightarrow \text{BaSO}_4\text{(s)} + 2 \text{NaCl(aq)}$

$$752 \text{ mg} = 0.752 \text{ g Na}_2\text{SO}_4 \times \frac{1 \text{ mol Na}_2\text{SO}_4}{142.1 \text{ g Na}_2\text{SO}_4} \times \frac{1 \text{ mol BaCl}_2}{1 \text{ mol Na}_2\text{SO}_4} \times \frac{1}{0.0558 \text{ L}} = 0.0948 \text{ M BaCl}_2$$

(d) $2 \text{HCl(aq)} + \text{Ca(OH)}_2\text{(aq)} \rightarrow \text{CaCl}_2\text{(aq)} + 2 \text{H}_2\text{O(l)}$

$$0.0427 \text{ L HCl} \times \frac{0.208 \text{ mol HCl}}{1 \text{ L HCl}} \times \frac{1 \text{ mol Ca(OH)}_2}{2 \text{ mol HCl}} \times \frac{74.10 \text{ g Ca(OH)}_2}{1 \text{ mol Ca(OH)}_2} = 0.329 \text{ g Ca(OH)}_2$$

4.83 *Analyze/Plan.* See Exercise 4.81(a) for a more detailed approach. *Solve.*

$$\frac{6.0 \text{ mol H}_2\text{SO}_4}{1 \text{ L}} \times 0.027 \text{ L} \times \frac{2 \text{ mol NaHCO}_3}{1 \text{ mol H}_2\text{SO}_4} \times \frac{84.01 \text{ g NaHCO}_3}{1 \text{ mol NaHCO}_3} = 27 \text{ g NaHCO}_3$$

4.84 See Exercise 4.81(a) for a more detailed approach.

$$\frac{0.115 \text{ mol NaOH}}{1 \text{ L}} \times 0.0425 \text{ L} \times \frac{1 \text{ mol CH}_3\text{COOH}}{1 \text{ mol NaOH}} \times \frac{60.05 \text{ g CH}_3\text{COOH}}{1 \text{ mol CH}_3\text{COOH}} = 0.29349 = 0.293 \text{ g CH}_3\text{COOH in 3.45 mL}$$

$$1.00 \text{ qt vinegar} \times \frac{1 \text{ L}}{1.057 \text{ qt}} \times \frac{1000 \text{ mL}}{1 \text{ L}} \times \frac{0.29349 \text{ g CH}_3\text{COOH}}{3.45 \text{ mL vinegar}} = 80.5 \text{ g CH}_3\text{COOH/qt}$$

4.85 *Analyze.* Given: M and vol HCl. Find: MM of base, an alkali metal hydroxide.

Plan. Alkali metal ions have a 1+ charge, so the general formula of an alkali metal hydroxide is MOH. One mol of MOH requires one mol of HCl for neutralization.

(a) $M \times L = \text{mol HCl} = \text{mol MOH}$. $\text{MM} = \frac{\text{g MOH}}{1 \text{ mol MOH}}$. *Solve.*

$$\frac{2.50 \text{ mol HCl}}{1 \text{ L}} \times 0.0170 \text{ L} = 0.0425 \text{ mol HCl} = 0.0425 \text{ mol MOH}$$

$$\text{MM of MOH} = \frac{4.36 \text{ g MOH}}{0.0425 \text{ mol MOH}} = 102.59 = 103 \text{ g/mol}$$

(b) MM of alkali metal = MM of MOH - (17.01 g) *Solve.*

$$\text{MM of alkali metal} = (102.59 \text{ g/mol} - 17.01 \text{ g/mol}) = 85.58 = 86 \text{ g/mol}$$

The experimental molar mass most closely fits that of Rb^+ , 85.47 g/mol

Check. The experimental molar mass matches one of the alkali metals.

4.86 *Analyze/Plan.* Follow the logic in Exercise 4.85. The unknown is a group 2A metal hydroxide, general formula $\text{M}(\text{OH})_2$. Two mol HCl are required to neutralize 1 mol $\text{M}(\text{OH})_2$. *Solve.*

(a) $\frac{2.50 \text{ mol HCl}}{1 \text{ L}} \times 0.0569 \text{ L} \times \frac{1 \text{ mol M}(\text{OH})_2}{2 \text{ mol HCl}} = 0.071125 = 0.0711 \text{ mol M}(\text{OH})_2$

$$\text{MM of M}(\text{OH})_2 = \frac{8.65 \text{ g M}(\text{OH})_2}{0.071125 \text{ mol M}(\text{OH})_2} = 121.62 = 122 \text{ g/mol}$$

(b) MM of group 2A metal = MM of $\text{M}(\text{OH})_2$ - 2(17.01 g)

$$\text{MM of group 2A metal} = (121.62 \text{ g/mol} - 34.02 \text{ g/mol}) = 87.60 = 88 \text{ g/mol}$$

The experimental molar mass most closely fits that of Sr^{2+} , 87.62 g/mol

Check. The experimental molar mass matches one of the group 2A metals.



(b) The precipitate is $\text{Ni}(\text{OH})_2$.

(c) *Plan.* Compare mol of each reactant; mol = $M \times L$

$$\text{Solve. } 0.200 \text{ M KOH} \times 0.1000 \text{ L KOH} = 0.0200 \text{ mol KOH}$$

$$0.150 \text{ M NiSO}_4 \times 0.2000 \text{ L NiSO}_4 = 0.0300 \text{ mol NiSO}_4$$

1 mol NiSO_4 requires 2 mol KOH, so 0.0300 mol NiSO_4 requires 0.0600 mol KOH. Because only 0.0200 mol KOH is available, KOH is the limiting reactant.

- (d) *Plan.* The amount of the limiting reactant (KOH) determines amount of product, in this case $\text{Ni}(\text{OH})_2$.

$$\text{Solve. } 0.0200 \text{ mol KOH} \times \frac{1 \text{ mol Ni}(\text{OH})_2}{2 \text{ mol KOH}} \times \frac{92.71 \text{ g Ni}(\text{OH})_2}{1 \text{ mol Ni}(\text{OH})_2} = 0.927 \text{ g Ni}(\text{OH})_2$$

- (e) *Plan/Solve.* Limiting reactant: OH^- : no excess OH^- remains in solution.

Excess reactant: Ni^{2+} : $M \text{ Ni}^{2+} \text{ remaining} = \text{mol Ni}^{2+} \text{ remaining} / \text{L solution}$

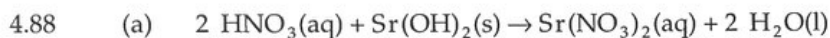
$$0.0300 \text{ mol Ni}^{2+} \text{ initial} - 0.0100 \text{ mol Ni}^{2+} \text{ reacted} = 0.0200 \text{ mol Ni}^{2+} \text{ remaining}$$

$$0.0200 \text{ mol Ni}^{2+} / 0.3000 \text{ L} = 0.0667 \text{ M Ni}^{2+}(\text{aq})$$

Spectators: SO_4^{2-} , K^+ . These ions do not react, so the only change in their concentration is dilution. The final volume of the solution is 0.3000 L.

$$M_2 = M_1 V_1 / V_2: 0.200 \text{ M K}^+ \times 0.1000 \text{ L} / 0.3000 \text{ L} = 0.0667 \text{ M K}^+(\text{aq})$$

$$0.150 \text{ M SO}_4^{2-} \times 0.2000 \text{ L} / 0.3000 \text{ L} = 0.100 \text{ M SO}_4^{2-}(\text{aq})$$



- (b) Determine the limiting reactant, then the identity and concentration of ions remaining in solution. Assume that the $\text{H}_2\text{O}(\text{l})$ produced by the reaction does **not** increase the total solution volume.

$$15.0 \text{ g Sr}(\text{OH})_2 \times \frac{1 \text{ mol Sr}(\text{OH})_2}{121.64 \text{ g Sr}(\text{OH})_2} = 0.1233 = 0.123 \text{ mol Sr}(\text{OH})_2$$

$$\text{mol OH}^- = 2(0.1233) \text{ mol Sr}(\text{OH})_2 = 0.2466 = 0.247 \text{ mol OH}^-$$

$$0.200 \text{ M HNO}_3 \times 0.0550 \text{ L HNO}_3 = 0.0110 \text{ mol HNO}_3.$$

Two mol HNO_3 react with one mol $\text{Sr}(\text{OH})_2$, so HNO_3 is the limiting reactant. No excess H^+ remains in solution. The remaining ions are OH^- (excess reactant), Sr^{2+} , and NO_3^- (spectators).

$$\text{OH}^-: 0.2466 \text{ mol OH}^- \text{ initial} - 0.0110 \text{ mol OH}^- \text{ react} = 0.2356 = 0.236 \text{ mol OH}^- \text{ remain}$$

$$0.2356 \text{ mol OH}^- / 0.0550 \text{ L soln} = 4.28 \text{ M OH}^-(\text{aq})$$

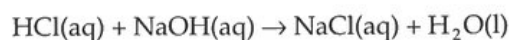
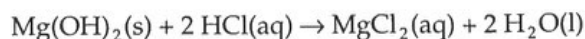
$$\text{Sr}^{2+}: 0.123 \text{ mol Sr}^{2+} / 0.0550 \text{ L soln} = 2.24 \text{ M Sr}^{2+}(\text{aq})$$

$$\text{NO}_3^-: 0.0110 \text{ mol NO}_3^- / 0.0550 \text{ L} = 0.200 \text{ M NO}_3^-(\text{aq})$$

- (c) The resulting solution is **basic** because of the large excess of $\text{OH}^-(\text{aq})$.



Find: mass % $\text{Mg}(\text{OH})_2$ in sample. *Plan/Solve.* Write balanced equations.



Calculate total moles $\text{HCl} = M \text{ HCl} \times L \text{ HCl}$

$$\frac{0.2050 \text{ mol HCl}}{1 \text{ L soln}} \times 0.1000 \text{ L} = 0.02050 \text{ mol HCl total}$$

- 4.57 (a) i. $\text{Zn(s)} + \text{Cd}^{2+}(\text{aq}) \rightarrow \text{Cd(s)} + \text{Zn}^{2+}(\text{aq})$
 ii. $\text{Cd(s)} + \text{Ni}^{2+}(\text{aq}) \rightarrow \text{Ni(s)} + \text{Cd}^{2+}(\text{aq})$

Observation (i) indicates that Cd is less active than Zn; observation (ii) indicates that Cd is more active than Ni. Cd is between Zn and Ni on the activity series.

- (b) Chromium, iron, and cobalt, the three elements between Zn and Ni in Table 4.5, more closely define the position of Cd in the activity series.
- (c) Place an iron strip in $\text{CdCl}_2(\text{aq})$. If Cd(s) is deposited, Cd is less active than Fe; if there is no reaction, Cd is more active than Fe. Do the same test with Co if Cd is less active than Fe or with Cr if Cd is more active than Fe.
- 4.58 $\text{Br}_2 + 2 \text{NaI} \rightarrow 2 \text{NaBr} + \text{I}_2$ indicates that Br_2 is more easily reduced than I_2 .
 $\text{Cl}_2 + 2 \text{NaBr} \rightarrow 2 \text{NaCl} + \text{Br}_2$ shows that Cl_2 is more easily reduced than Br_2 .
 The order for ease of reduction is $\text{Cl}_2 > \text{Br}_2 > \text{I}_2$. Conversely, the order for ease of oxidation is $\text{I}^- > \text{Br}^- > \text{Cl}^-$.
- (a) From the information above, the halogen I_2 is most stable (less likely to react) when mixed with other halides, X^- .
- (b) $\text{Cl}_2 + 2 \text{KI} \rightarrow 2 \text{KCl} + \text{I}_2$
- (c) $\text{Br}_2 + \text{LiCl} \rightarrow$ no reaction

Concentrations of Solutions (Section 4.5)

- 4.59 (a) *Concentration* is an *intensive* property; it is the **ratio** of the amount of solute present in a certain quantity of solvent or solution. This ratio remains constant regardless of how much solution is present.
- (b) The term *0.50 mol HCl* defines an amount (~18 g) of the pure substance HCl. The term *0.50 M HCl* is a ratio; it indicates that there are 0.50 mol of HCl solute in 1.0 liter of solution. This same ratio of moles solute to solution volume is present regardless of the volume of solution under consideration.
- 4.60 *Analyze/Plan.* Follow the logic in Sample Exercise 4.11. *Solve.*
- (a) $M = \frac{\text{mol solute}}{\text{L solution}}; \frac{35.0 \text{ g C}_{12}\text{H}_{22}\text{O}_{11}}{1.000 \text{ L}} \times \frac{1 \text{ mol C}_{12}\text{H}_{22}\text{O}_{11}}{342.3 \text{ g C}_{12}\text{H}_{22}\text{O}_{11}} = 0.102 \text{ M C}_{12}\text{H}_{22}\text{O}_{11}$
- (b) Add 1.000 L of water to reduce the molarity by a factor of 2. Adding water does not change to amount of solute, but it does change the total volume of solution. A total solution volume of 2.000 L will reduce the molarity by a factor of 2.
- 4.61 *Analyze/Plan.* Follow the logic in Sample Exercises 4.11 and 4.12. *Solve.*
- (a) $M = \frac{\text{mol solute}}{\text{L solution}}; \frac{0.175 \text{ mol ZnCl}_2}{150 \text{ mL}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 1.17 \text{ M ZnCl}_2$
- Check.* $(0.175 / 0.150) > 1.0 \text{ M}$

$$(b) \quad \text{mol} = M \times L; \frac{4.50 \text{ mol HNO}_3}{1 \text{ L}} \times \frac{1 \text{ mol H}^+}{\text{mol HNO}_3} \times 0.0350 \text{ L} = 0.158 \text{ mol H}^+$$

Check. $(4.5 \times .04) \approx 0.16 \text{ mol}$

$$(c) \quad L = \frac{\text{mol}}{M}; \frac{0.350 \text{ mol NaOH}}{6.00 \text{ mol NaOH/L}} = 0.0583 \text{ L or } 58.3 \text{ mL of } 6.00 \text{ M NaOH}$$

Check. $(0.325/6.0) > 0.50 \text{ L}$.

$$4.62 \quad (a) \quad M = \frac{\text{mol solute}}{\text{L solution}}; \frac{12.5 \text{ g Na}_2\text{CrO}_4}{0.750 \text{ L}} \times \frac{1 \text{ mol Na}_2\text{CrO}_4}{161.97 \text{ g Na}_2\text{CrO}_4} = 0.103 \text{ M Na}_2\text{CrO}_4$$

$$(b) \quad \text{mol} = M \times L; \frac{0.112 \text{ mol KBr}}{1 \text{ L}} \times 0.150 \text{ L} = 1.68 \times 10^{-2} \text{ mol KBr}$$

$$(c) \quad L = \frac{\text{mol}}{M}; \frac{0.150 \text{ mol HCl}}{6.1 \text{ mol HCl/L}} = 2.5 \times 10^{-2} \text{ L or } 25 \text{ mL}$$

4.63 *Analyze.* Given molarity, M , and volume, L , find mass of Na^+ (aq) in the blood.

Plan. Calculate moles Na^+ (aq) using the definition of molarity: $M = \frac{\text{mol}}{\text{L}}$; $\text{mol} = M \times L$.

Calculate mass Na^+ (aq) using the definition moles: $\text{mol} = \text{g}/\text{MM}$; $\text{g} = \text{mol} \times \text{MM}$. (MM is the symbol for molar mass in this manual.)

$$\text{Solve. } \frac{0.135 \text{ mol}}{\text{L}} \times 5.0 \text{ L} \times \frac{23.0 \text{ g Na}^+}{\text{mol Na}^+} = 15.525 = 16 \text{ g Na}^+ \text{ (aq)}$$

Check. Because there are more than 0.1 mol/L and we have 5.0 L, there should be more than half a mol (11.5 g) of Na^+ . The calculation agrees with this estimate.

4.64 Calculate the mol of Na^+ at the two concentrations; the difference is the mol NaCl required to increase the Na^+ concentration to the desired level.

$$\frac{0.118 \text{ mol}}{\text{L}} \times 4.6 \text{ L} = 0.5428 = 0.54 \text{ mol Na}^+$$

$$\frac{0.138 \text{ mol}}{\text{L}} \times 4.6 \text{ L} = 0.6348 = 0.63 \text{ mol Na}^+$$

$(0.6348 - 0.5428) = 0.092 = 0.09 \text{ mol NaCl}$ (2 decimal places and 1 sig fig)

$$0.092 \text{ mol NaCl} \times \frac{58.5 \text{ g NaCl}}{\text{mol}} = 5.38 = 5 \text{ g NaCl}$$

4.65 *Analyze.* Given: g alcohol/100 mL blood; molecular formula of alcohol. Find: molarity (mol/L) of alcohol. *Plan.* Use the molar mass (MM) of alcohol to change (g/100) mL to (mol/100 mL) then mL to L.

Solve. MM of alcohol = $2(12.01) + 6(1.1008) + 1(16.00) = 46.07 \text{ g alcohol/mol}$

$$\text{BAC} = \frac{0.08 \text{ g alcohol}}{100 \text{ mL blood}} \times \frac{1 \text{ mol alcohol}}{46.07 \text{ g alcohol}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 0.0174 = 0.02 \text{ M alcohol}$$