4.71 How would you prepare 60.0 mL of 0.200 M HNO₃ from a stock solution of 4.00 M HNO₃?

\[ M_{\text{initial}}V_{\text{initial}} = M_{\text{final}}V_{\text{final}} \]

\[ V_{\text{initial}} = \frac{M_{\text{final}} \times V_{\text{final}}}{M_{\text{initial}}} = \frac{0.200 \times 60.00 \text{ mL}}{4.00 \text{ M}} = 3.00 \text{ mL} \]

To prepare the 0.200 M solution, you would dilute 3.00 mL of the 4.00 M HNO₃ solution to a final volume of 60.0 mL.

4.99 Classify the following reactions as one of the following: precipitation, acid-base, combination redox, decomposition redox, disproportionation redox, or displacement redox

(a) Cl₂ + 2 OH⁻ → Cl⁻ + ClO⁻ + H₂O  disproportionation redox reaction
(b) Ca²⁺ + CO₃²⁻ → CaCO₃  precipitation
(c) NH₃ + H⁺ → NH₄⁺  acid-base reaction
(d) 2 CCl₄ + CrO₄²⁻ → 2 COCl₂ + Cr₂O₇²⁻ + 2 Cl⁻  Is this a redox reaction?  no  (Don’t try to more specifically categorize this reaction.)
(e) Ca + F₂ → CaF₂  Combination redox reaction.
(f) 2 Li + H₂ → 2 LiH:  Combination Redox
(g) Ba(NO₃)₂ + Na₂SO₄ → 2 NaNO₃ + BaSO₄:  Precipitation
(h) CuO + H₂ → Cu + H₂O:  Displacement Redox
(i) Zn + 2 HCl → ZnCl₂ + H₂:  Displacement Redox
(j) 2 FeCl₂ + Cl₂ → 2 FeCl₃:  Combination Redox

4.101 Which of the following aqueous solutions would you expect to be the best conductor of electricity at 25°C? Explain your answer.

a) 0.20 M NaCl  
   (0.20 M NaCl)(2 ions/mole NaCl) = 0.40 M for ions

b) 0.60 M CH₃COOH  
   Acetic acid is a weak acid. Much less than 0.60M of ions will be in soln

c) 0.25 M HCl  
   Strong acid so (0.25 M HCl)(2 ions/mole of HCl) = 0.50 M for ions

(d) 0.20 M Mg(NO₃)₂,  
   [0.20 M Mg(NO₃)₂][3 moles ions/mole Mg(NO₃)₂] = 0.60M for ions
4.102 A 5.00 \times 10^2$-mL sample of 2.00 $M$ HCl solution is treated with 4.47 g of magnesium. Calculate the concentration of the acid solution after all the metal has reacted. Assume that the volume remains unchanged.

\[
\text{Molecular Eq: } \text{Mg}(s) + 2\text{HCl}(aq) \rightarrow \text{MgCl}_2(aq) + \text{H}_2(g)
\]

4.47 g Mg \times \frac{1 \text{ mol Mg}}{24.31 \text{ g Mg}} \times \frac{2 \text{ mol HCl}}{1 \text{ mol Mg}} = 0.368 \text{ mol HCl reacted}

\[5.00 \times 10^2 \text{ mL} \times \frac{2.00 \text{ mol HCl}}{1000 \text{ mL soln}} = 1.00 \text{ mol HCl in original solution}\]

1.00 mol HCl originally – 0.368 mol HCl reacted = 0.632 mol HCl remaining

\[\text{conc. of HCl after reaction} = \frac{0.632 \text{ mol HCl}}{0.500 \text{ L}} = 1.26 \text{ mol/L} = \boxed{1.26 \text{ M}}\]

4.109 Calculate the mass of the precipitate formed when 2.27 L of 0.0820 $M$ Ba(OH)$_2$ are mixed with 3.07 L of 0.0774 $M$ Na$_2$SO$_4$.

\[
\text{Molecular Eq: } \text{Ba(OH)}_2(aq) + \text{Na}_2\text{SO}_4(aq) \rightarrow \text{BaSO}_4(s) + 2\text{NaOH}(aq)
\]

moles Ba(OH)$_2$: $\frac{(2.27 \text{ L})(0.0820 \text{ mol/L})}{1 \text{ L}} = 0.186$ mol Ba(OH)$_2$

moles Na$_2$SO$_4$: $\frac{(3.07 \text{ L})(0.0664 \text{ mol/L})}{1 \text{ L}} = 0.203$ mol Na$_2$SO$_4$

Since the mole ratio between Ba(OH)$_2$ and Na$_2$SO$_4$ is 1:1, so \textbf{Ba(OH)$_2$ is the limiting reagent.}

\[
0.186 \text{ mol Ba(OH)$_2$} \times \frac{1 \text{ mol BaSO}_4}{1 \text{ mol Ba(OH)$_2$}} \times \frac{233.4 \text{ g BaSO}_4}{1 \text{ mol BaSO}_4} = 43.4 \text{ g BaSO}_4
\]

4.112 A 1.00-g sample of a metal X (that is known to form $X^{2+}$ ions) was added to a 0.100 L of 0.500 $M$ H$_2$SO$_4$. After all the metal had reacted, the remaining acid required 0.0334 L of 0.500 $M$ NaOH solution for neutralization. Calculate the molar mass of the metal and identify the element.

\[
\text{Mol Eq for 1$^{\text{st}}$ rxn: } \text{X}(s) + \text{H}_2\text{SO}_4(aq) \rightarrow \text{XSO}_4(aq) + \text{H}_2(g)
\]

\[
\text{Mol Eq for 2$^{\text{nd}}$ rxn: } \text{H}_2\text{SO}_4(aq) + 2\text{NaOH}(aq) \rightarrow \text{Na}_2\text{SO}_4(aq) + 2\text{H}_2\text{O}(l)
\]

First, let’s find the number of moles of excess acid from the reaction with NaOH.

\[
0.0334 \text{ L} \times \frac{0.500 \text{ mol NaOH}}{1 \text{ L soln}} \times \frac{1 \text{ mol H}_2\text{SO}_4}{2 \text{ mol NaOH}} = 8.35 \times 10^{-3} \text{ mol H}_2\text{SO}_4 \text{ (excess acid leftover after rxn \#1)}
\]

\[
0.100 \text{ L} \times \frac{0.500 \text{ mol H}_2\text{SO}_4}{1 \text{ L soln}} = 0.0500 \text{ mol H}_2\text{SO}_4 \text{ originally added to X (s)}
\]

\[(0.0500 \text{ mol H}_2\text{SO}_4) – (8.35 \times 10^{-3} \text{ mol H}_2\text{SO}_4) = 0.0417 \text{ mol H}_2\text{SO}_4 \text{ that reacted with metal, X}
\]

Since the mole ratio from the balanced equation is 1 mole X : 1 mole H$_2$SO$_4$, then the amount of X that reacted is 0.0417 mol X.

\[
\text{molar mass X} = \frac{1.00 \text{ g X}}{0.0417 \text{ mol X}} = \boxed{24.0 \text{ g/mol}} \text{ The element is magnesium.}
\]
4.114 A 60.0-mL 0.513 M glucose (C₆H₁₂O₆) solution is mixed with 120.0 mL of 2.33 M glucose solution. What is the concentration of the final solution? Assume the volumes are additive. First, calculate the number of moles of glucose present.

\[
\frac{0.513 \text{ mol glucose}}{1000 \text{ mL soln}} \times 60.0 \text{ mL} = 0.0308 \text{ mol glucose}
\]

\[
\frac{2.33 \text{ mol glucose}}{1000 \text{ mL soln}} \times 120.0 \text{ mL} = 0.280 \text{ mol glucose}
\]

Total Volume = 60.0 mL + 120.0 mL = 180.0 mL = 0.180 L

**Molarity of final solution** = \(\frac{(0.0308 + 0.280) \text{ mol glucose}}{0.180 \text{ L}} = 1.73 \text{ mol/L} = 1.73 \text{ M}\)

4.118 Using the apparatus shown at the right, a student found that a sulfuric acid solution caused the light bulb to glow brightly. However, after the addition of a certain amount of barium hydroxide [Ba(OH)₂] solution, the light began to dim even though Ba(OH)₂ is also a strong electrolyte. Determine the molecular and net ionic equations for the reaction that occurs and explain why the light bulb dims.

**Molecular Eq:** \(H_2SO_4(aq) + Ba(OH)_2(aq) \rightarrow 2H_2O(l) + BaSO_4(s)\)

**Net Ionic:** \(2H^+(aq) + SO_4^{2-}(aq) + Ba^{2+}(aq) + 2OH^-(aq) \rightarrow 2H_2O(l) + BaSO_4(s)\)

\(H^+\) from the sulfuric acid reacts with the \(OH^-\) from the barium hydroxide to form water. The barium cations react with the sulfate anions to form insoluble barium sulfate. Thus, the reaction depletes the solution of ions and the conductivity decreases.

4.121 The concentration of lead ions (Pb²⁺) in a sample of polluted water that also contains NO₃⁻ ions is determined by adding solid Na₂SO₄ to exactly 500. mL of the water.

(a) Write the molecular and net ionic equation for the reaction.

**Mol Eq:** \(\text{Pb(NO}_3\text{)}_2(aq) + \text{Na}_2\text{SO}_4(aq) \rightarrow \text{PbSO}_4(s) + 2\text{NaNO}_3(aq)\)

**Net Ionic:** \(\text{Pb}^{2+}(aq) + \text{SO}_4^{2-}(aq) \rightarrow \text{PbSO}_4(s)\)

(b) Calculate the molar concentration of Pb²⁺ if 0.00450 g of Na₂SO₄ was needed for the complete precipitation of Pb²⁺ as PbSO₄.

\[
0.00450 \text{ g Na}_2\text{SO}_4 \times \frac{1 \text{ mol Na}_2\text{SO}_4}{142.1 \text{ g Na}_2\text{SO}_4} \times \frac{1 \text{ mol Pb(NO}_3\text{)}_2}{1 \text{ mol Na}_2\text{SO}_4} \times \frac{1 \text{ mol Pb}^{2+}}{1 \text{ mol Pb(NO}_3\text{)}_2} = 3.17 \times 10^{-5} \text{ mol Pb}^{2+}
\]

\[
\frac{[\text{Pb}^{2+}]}{\text{L of soln}} = \frac{3.17 \times 10^{-5} \text{ mol Pb}^{2+}}{0.500 \text{ L soln}} = 6.34 \times 10^{-5} \text{ M}
\]
4.132 A useful application of oxalic acid is the removal of rust (Fe₂O₃) from bathtub rings according to the reaction,  

\[ \text{Fe}_2\text{O}_3(s) + 6\text{H}_2\text{C}_2\text{O}_4(aq) \rightarrow 2\text{Fe(C}_2\text{O}_4)_3^{3+} + 3\text{H}_2\text{O} + 6\text{H}^+(aq) \]

Calculate the number of grams of rust that can be removed by 5.00 \times 10^2 \text{ mL} of a 0.100 \text{ M solution of oxalic acid.}

\[
5.00 \times 10^2 \text{ mL} \times \frac{0.100 \text{ mol H}_2\text{C}_2\text{O}_4}{1000 \text{ mL soln}} = 0.0500 \text{ mol H}_2\text{C}_2\text{O}_4
\]

\[
0.0500 \text{ mol H}_2\text{C}_2\text{O}_4 \times \frac{1 \text{ mol Fe}_2\text{O}_3}{6 \text{ mol H}_2\text{C}_2\text{O}_4} \times \frac{159.7 \text{ g Fe}_2\text{O}_3}{1 \text{ mol Fe}_2\text{O}_3} = \underline{1.33 \text{ g Fe}_2\text{O}_3}
\]

A. You perform a Beer’s law experiment to determine the concentration of an unknown solution of copper (II) ion in solution. You use a spectrometer and collect absorbance of 4 stock solution at a wavelength of 635 nm. The data is displayed below at the right.

1. Sketch the data in the graph below the data chart.

2. Write a linear equation that relates the concentration of CuSO₄ to its absorbance at 635 nm. The value for the slope of the line must have units.

\[ \text{A} = \text{abC} \text{ where the slope = ab} \]

\[
\text{ab} = \frac{0.960 - 0.240}{100\text{mM} - 25\text{mM}} = \frac{0.720}{75.0\text{mM}} = 0.00960 \text{ mM}^{-1}
\]

Thus, \[ \text{A} = (0.00960 \text{ mM}^{-1})(C) \]

3. You now put a 2.50 g penny into 50.0 mL of HNO₃. After reacting overnight, you dilute the blue Cu²⁺ solution to a volume of 100.0 mL. You test a sample of this solution and find that it has an absorbance of 0.120 at 635 nm.

a. Calculate the concentration of Cu²⁺ in the solution.

\[ 0.120 = (0.00960 \text{ mM}^{-1})(C) \]

\[ C = \frac{0.120}{0.00960 \text{ mM}^{-1}} = 12.5 \text{ mM} \]

b. Calculate the mass (in grams) of copper present in the penny.

\[
0.1000\text{L} \times \frac{12.5 \text{ mmol Cu}^{2+}}{1 \text{ L}} \times \frac{1 \text{ mol}}{1000 \text{ mmol}} = 0.00125 \text{ mol Cu}^{2+} \times \frac{63.5 \text{ g Cu}^{2+}}{1 \text{ mol}} = \underline{0.0793 \text{ g Cu}}
\]

c. Calculate the percent of copper present in the penny.

\[ \% \text{ Cu} = \frac{0.0793 \text{ g Cu}}{2.50 \text{ g penny}} \times 100 = 3.18\% \]