Chapter 4 Suggested end-of-chapter problems with solutions

27. Calculate the molarity of each of these solutions.

a. A **5.623-g** sample of $\mathbf{NaHCO_3}$ is dissolved in enough water to make **250.0 mL** of solution.

b. A **184.6-mg** sample of $\mathbf{K_2Cr_2O_7}$ is dissolved in enough water to make **500.0 mL** of solution.

c. A **0.1025-g** sample of copper metal is dissolved in **35 mL** of concentrated $\mathbf{HNO_3}$ to form $\mathbf{Cu^{2+}}$ ions and then water is added to make a total volume of **200.0 mL** . (Calculate the molarity of $\mathbf{Cu^{2+}}$.)

Answer:

a. $$5.623 \text{ g NaHCO}_3 \times \frac{1 \text{ mol NaHCO}_3}{84.01 \text{ g NaHCO}_3} = 6.693 \times 10^{-2} \text{ mol NaHCO}_3$$

$$M = \frac{6.693 \times 10^{-2} \text{ mol}}{250.0 \text{ mL}} \times \frac{1000 \text{ mL}}{\text{L}} = 0.2677\ M \text{ NaHCO}_3$$

b. $$0.1846 \text{ g K}_2\text{Cr}_2\text{O}_7 \times \frac{1 \text{ mol K}_2\text{Cr}_2\text{O}_7}{294.20 \text{ g K}_2\text{Cr}_2\text{O}_7} = 6.275 \times 10^{-4} \text{ mol K}_2\text{Cr}_2\text{O}_7$$

$$M = \frac{6.275 \times 10^{-4} \text{ mol}}{500.0 \times 10^{-3} \text{ L}} = 1.255 \times 10^{-3}\ M \text{ K}_2\text{Cr}_2\text{O}_7$$

c. $$0.1025 \text{ g Cu} \times \frac{1 \text{ mol Cu}}{63.55 \text{ g Cu}} = 1.613 \times 10^{-3} \text{ mol Cu} = 1.613 \times 10^{-3} \text{ mol Cu}^{2+}$$

$$M = \frac{1.613 \times 10^{-3} \text{ mol Cu}^{2+}}{200.0 \text{ mL}} \times \frac{1000 \text{ mL}}{\text{L}} = 8.065 \times 10^{-3}\ M \text{ Cu}^{2+}$$

28. A solution of ethanol (**C_2H_5OH**) in water is prepared by dissolving **75.0 mL** of ethanol $\left(\textbf{density} = \mathbf{0.79\ g/cm^3}\right)$ in enough water to make **250.0 mL** of solution. What is the molarity of the ethanol in this solution?

Answer: $75.0\text{ mL} \times \dfrac{0.79\text{ g}}{\text{mL}} \times \dfrac{1\text{ mol}}{46.07\text{ g}} = 1.3\text{ mol } C_2H_5OH;\ \text{molarity} = \dfrac{1.3\text{ mol}}{0.250\text{ L}} = 5.2\ M\ C_2H_5OH$

37. A solution is prepared by dissolving **10.8 g** ammonium sulfate in enough water to make **100.0 mL** of stock solution. A **10.00-mL** sample of this stock solution is added to **50.00 mL** of water. Calculate the concentration of ammonium ions and sulfate ions in the final solution.

Answer: $10.8\text{ g } (NH_4)_2SO_4 \times \dfrac{1\text{ mol}}{132.15\text{ g}} = 8.17 \times 10^{-2}\text{ mol } (NH_4)_2SO_4$

$$\text{Molarity} = \frac{8.17 \times 10^{-2}\text{ mol}}{100.0\text{ mL}} \times \frac{1000\text{ mL}}{\text{L}} = 0.817\ M\ (NH_4)_2SO_4$$

Moles of $(NH_4)_2SO_4$ in final solution:

$$10.00 \times 10^{-3}\text{ L} \times \frac{0.817\text{ mol}}{\text{L}} = 8.17 \times 10^{-3}\text{ mol}$$

$$\text{Molarity of final solution} = \frac{8.17 \times 10^{-3}\text{ mol}}{(10.00 + 50.00)\text{ mL}} \times \frac{1000\text{ mL}}{\text{L}} = 0.136\ M\ (NH_4)_2SO_4$$

$$(NH_4)_2SO_4(s) \rightarrow 2\ NH_4^+(aq) + SO_4^{2-}(aq);\ \ M_{NH_4^+} = 2(0.136) = 0.272\ M;\ \ M_{SO_4^{2-}} = 0.136\ M$$

38. A solution was prepared by mixing **50.00 mL** of **0.100 *M* HNO_3** and **100.00 mL** of **0.200 *M* HNO_3** . Calculate the molarity of the final solution of nitric acid.

Answer: $\text{Molarity} = \dfrac{\text{total mol } HNO_3}{\text{total volume}}$; total volume = 0.05000 L + 0.10000 L = 0.15000 L

$$\text{Total mol } HNO_3 = 0.05000 \text{ L} \times \frac{0.100 \text{ mol } HNO_3}{\text{L}} + 0.10000 \text{ L} \times \frac{0.200 \text{ mol } HNO_3}{\text{L}}$$

$$\text{Total mol } HNO_3 = 5.00 \times 10^{-3} \text{ mol} + 2.00 \times 10^{-2} \text{ mol} = 2.50 \times 10^{-2} \text{ mol } HNO_3$$

$$\text{Molarity} = \frac{2.50 \times 10^{-2} \text{ mol } HNO_3}{0.15000 \text{ L}} = 0.167\ M\ HNO_3$$

As expected, the molarity of HNO_3 is between 0.100 *M* and 0.200 *M*.

43. On the basis of the general solubility rules given in Table 4.1, predict which of the following substances are likely to be soluble in water.

a. aluminum nitrate

b. magnesium chloride

c. rubidium sulfate

d. nickel(II) hydroxide

e. lead(II) sulfide

f. magnesium hydroxide

g. iron(III) phosphate

Answer: The solubility rules referenced in the following answers are outlined in Table 4.1 of the text.

a. Soluble: Most nitrate salts are soluble (Rule 1).

b. Soluble: Most chloride salts are soluble except for Ag^+, Pb^{2+}, and Hg_2^{2+} (Rule 3).

c. Soluble: Most sulfate salts are soluble except for $BaSO_4$, $PbSO_4$, Hg_2SO_4, and $CaSO_4$ (Rule 4.)

d. Insoluble: Most hydroxide salts are only slightly soluble (Rule 5).

Note: We will interpret the phrase "slightly soluble" as meaning insoluble and the phrase "marginally soluble" as meaning soluble. So the marginally soluble hydroxides $Ba(OH)_2$, $Sr(OH)_2$, and $Ca(OH)_2$ will be assumed soluble unless noted otherwise.

e. Insoluble: Most sulfide salts are only slightly soluble (Rule 6). Again, "slightly soluble" is interpreted as "insoluble" in problems like these.

f. Insoluble: Rule 5 (see answer d).

g. Insoluble: Most phosphate salts are only slightly soluble (Rule 6).

45. When the following solutions are mixed together, what precipitate (if any) will form?

a. $\mathbf{FeSO_4}(\boldsymbol{aq}) + \mathbf{KCl}(\boldsymbol{aq})$

b. $\mathbf{Al(NO_3)_3}(\boldsymbol{aq}) + \mathbf{Ba(OH)_2}(\boldsymbol{aq})$

c. $\mathbf{CaCl_2}(\boldsymbol{aq}) + \mathbf{Na_2SO_4}(\boldsymbol{aq})$

d. $\mathbf{K_2S}(\boldsymbol{aq}) + \mathbf{Ni(NO_3)_2}(\boldsymbol{aq})$

Answer: In these reactions, soluble ionic compounds are mixed together. To predict the precipitate, switch the anions and cations in the two reactant compounds to predict possible products; then use the solubility rules in Table 4.1 to predict if any of these possible products are insoluble (are the precipitate). Note that the phrase "slightly soluble" in Table 4.1 is interpreted to mean insoluble, and the phrase "marginally soluble" is interpreted to mean soluble.

a. Possible products = $FeCl_2$ and K_2SO_4; both salts are soluble, so no precipitate forms.

b. Possible products = $Al(OH)_3$ and $Ba(NO_3)_2$; precipitate = $Al(OH)_3(s)$

c. Possible products = $CaSO_4$ and NaCl; precipitate = $CaSO_4(s)$

d. Possible products = KNO_3 and NiS; precipitate = NiS(s)

51. Write net ionic equations for the reaction, if any, that occurs when aqueous solutions of the following are mixed.

a. ammonium sulfate and barium nitrate

b. lead(II) nitrate and sodium chloride

c. sodium phosphate and potassium nitrate

d. sodium bromide and rubidium chloride

e. copper(II) chloride and sodium hydroxide

Answer:

a. $(NH_4)_2SO_4(aq) + Ba(NO_3)_2(aq) \rightarrow 2\ NH_4NO_3(aq) + BaSO_4(s)$

$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$

b. $Pb(NO_3)_2(aq) + 2\ NaCl(aq) \rightarrow PbCl_2(s) + 2\ NaNO_3(aq)$

$Pb^{2+}(aq) + 2\ Cl^-(aq) \rightarrow PbCl_2(s)$

c. Potassium phosphate and sodium nitrate are both soluble in water. No reaction occurs.

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

e. $CuCl_2(aq) + 2\ NaOH(aq) \rightarrow Cu(OH)_2(s) + 2\ NaCl(aq)$

$Cu^{2+}(aq) + 2\ OH^-(aq) \rightarrow Cu(OH)_2(s)$

54. A sample may contain any or all of the following ions: $\mathbf{Hg_2^{2+}}$, $\mathbf{Ba^{2+}}$, and $\mathbf{Mn^{2+}}$.

a. No precipitate formed when an aqueous solution of **NaCl** was added to the sample solution.

b. No precipitate formed when an aqueous solution of $\mathbf{Na_2SO_4}$ was added to the sample solution.

c. A precipitate formed when the sample solution was made basic with **NaOH** .

Which ion or ions are present in the sample solution?

Answer: Because no precipitates formed upon addition of NaCl or Na_2SO_4, we can conclude that Hg_2^{2+} and Ba^{2+} are not present in the sample because Hg_2Cl_2 and $BaSO_4$ are insoluble salts. However, Mn^{2+} may be present since Mn^{2+} does not form a precipitate with either NaCl or Na_2SO_4. A precipitate formed with NaOH; the solution must contain Mn^{2+} because it forms a precipitate with OH^- [$Mn(OH)_2(s)$].

59. What mass of solid **AgBr** is produced when **100.0 mL** of **0.150 *M* $\mathbf{AgNO_3}$** is added to **20.0 mL** of **1.00 *M* NaBr** ?

Answer: The reaction is $AgNO_3(aq) + NaBr(aq) \rightarrow AgBr(s) + NaNO_3(aq)$.

Assuming $AgNO_3$ is limiting:

$$100.0 \text{ mL } AgNO_3 \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.150 \text{ mol } AgNO_3}{\text{L } AgNO_3} \times \frac{1 \text{ mol AgBr}}{\text{mol } AgNO_3} \times \frac{187.8 \text{ g AgBr}}{\text{mol AgBr}}$$

$$= 2.82 \text{ g AgBr}$$

Assuming NaBr is limiting:

$$20.0 \text{ mL NaBr} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{1.00 \text{ mol NaBr}}{\text{L NaBr}} \times \frac{1 \text{ mol AgBr}}{\text{mol NaBr}} \times \frac{187.8 \text{ g AgBr}}{\text{mol AgBr}}$$

$$= 3.76 \text{ g AgBr}$$

The $AgNO_3$ reagent produces the smaller quantity of AgBr, so $AgNO_3$ is limiting and 2.82 g AgBr can form.

63. A **1.42-g** sample of a pure compound, with formula $\mathbf{M_2SO_4}$, was dissolved in water and treated with an excess of aqueous calcium chloride, resulting in the precipitation of all the sul-fate ions as calcium sulfate. The precipitate was collected, dried, and found to weigh **1.36 g** . Determine the atomic mass of **M** , and identify **M** .

Answer: $M_2SO_4(aq) + CaCl_2(aq) \rightarrow CaSO_4(s) + 2\ MCl(aq)$

$$1.36\text{ g } CaSO_4 \times \frac{1\text{ mol } CaSO_4}{136.15\text{ g } CaSO_4} \times \frac{1\text{ mol } M_2SO_4}{\text{mol } CaSO_4} = 9.99 \times 10^{-3}\text{ mol } M_2SO_4$$

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

$$\text{molar mass} = \frac{1.42\text{ g } M_2SO_4}{9.99 \times 10^{-3}\text{ mol } M_2SO_4} = 142\text{ g/mol}$$

142 u = 2(atomic mass M) + 32.07 + 4(16.00), atomic mass M = 23 u

From periodic table, M = Na (sodium).

65. Write the balanced formula, complete ionic, and net ionic equations for each of the following acid–base reactions.

a. $\mathbf{HClO_4}(aq) + \mathbf{Mg(OH)_2}(s) \rightarrow$

b. $\mathbf{HCN}(aq) + \mathbf{NaOH}(aq) \rightarrow$

c. $\mathbf{HCl}(aq) + \mathbf{NaOH}(aq) \rightarrow$

Answer: All the bases in this problem are ionic compounds containing OH^-. The acids are either strong or weak electrolytes. The best way to determine if an acid is a strong or weak electrolyte is to memorize all the strong electrolytes (strong acids). Any other acid you encounter that is not a strong acid will be a weak electrolyte (a weak acid), and the formula should be left unaltered in the complete ionic and net ionic equations. The strong acids to recognize are HCl, HBr, HI, HNO_3, $HClO_4$, and H_2SO_4. For the following answers, the order of the equations are formula, complete ionic, and net ionic.

a. $2\ HClO_4(aq) + Mg(OH)_2(s) \rightarrow 2\ H_2O(l) + Mg(ClO_4)_2(aq)$

$2\ H^+(aq) + 2\ ClO_4^-(aq) + Mg(OH)_2(s) \rightarrow 2\ H_2O(l) + Mg^{2+}(aq) + 2\ ClO_4^-(aq)$

$2\ H^+(aq) + Mg(OH)_2(s) \rightarrow 2\ H_2O(l) + Mg^{2+}(aq)$

b. $HCN(aq) + NaOH(aq) \rightarrow H_2O(l) + NaCN(aq)$

$HCN(aq) + Na^+(aq) + OH^-(aq) \rightarrow H_2O(l) + Na^+(aq) + CN^-(aq)$

$HCN(aq) + OH^-(aq) \rightarrow H_2O(l) + CN^-(aq)$

c. $HCl(aq) + NaOH(aq) \rightarrow H_2O(l) + NaCl(aq)$

$H^+(aq) + Cl^-(aq) + Na^+(aq) + OH^-(aq) \rightarrow H_2O(l) + Na^+(aq) + Cl^-(aq)$

$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

68. What acid and what base would react in aqueous solution so that the following salts appear as products in the formula equation? Write the balanced formula equation for each reaction.

a. potassium perchlorate

b. cesium nitrate

c. calcium iodide

Answer:

a. Perchloric acid plus potassium hydroxide is a possibility.

$HClO_4(aq) + KOH(aq) \rightarrow H_2O(l) + KClO_4(aq)$

b. Nitric acid plus cesium hydroxide is a possibility.

$HNO_3(aq) + CsOH(aq) \rightarrow H_2O(l) + CsNO_3(aq)$

c. Hydroiodic acid plus calcium hydroxide is a possibility.

$2\ HI(aq) + Ca(OH)_2(aq) \rightarrow 2\ H_2O(l) + CaI_2(aq)$

69. What volume of each of the following acids will react completely with **50.00 mL** of **0.200 *M* NaOH** ?

a. **0.100 *M* HCl**

b. **0.150 *M*** $\mathbf{HNO_3}$

c. **0.200 *M*** $\mathbf{HC_2H_3O_2}$ (**1** acidic hydrogen)

Answer:

If we begin with 50.00 mL of 0.200 *M* NaOH, then:

$$50.00 \times 10^{-3}\ \text{L} \times \frac{0.200\ \text{mol}}{\text{L}} = 1.00 \times 10^{-2}\ \text{mol NaOH is to be neutralized}$$

a. $NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$

$$1.00 \times 10^{-2}\ \text{mol NaOH} \times \frac{1\ \text{mol HCl}}{\text{mol NaOH}} \times \frac{1\ \text{L}}{0.100\ \text{mol}} = 0.100\ \text{L or 100. mL}$$

b. $HNO_3(aq) + NaOH(aq) \rightarrow H_2O(l) + NaNO_3(aq)$

$$1.00 \times 10^{-2}\ \text{mol NaOH} \times \frac{1\ \text{mol HNO}_3}{\text{mol NaOH}} \times \frac{1\ \text{L}}{0.150\ \text{mol HNO}_3} = 6.67 \times 10^{-2}\ \text{L or 66.7 mL}$$

c. $HC_2H_3O_2(aq) + NaOH(aq) \rightarrow H_2O(l) + NaC_2H_3O_2(aq)$

$$1.00 \times 10^{-2}\ \text{mol NaOH} \times \frac{1\ \text{mol HC}_2\text{H}_3\text{O}_2}{\text{mol NaOH}} \times \frac{1\ \text{L}}{0.200\ \text{mol HC}_2\text{H}_3\text{O}_2} = 5.00 \times 10^{-2}\ \text{L}$$

71. Hydrochloric acid **(75.0 mL of 0.250 M)** is added to **225.0 mL** of **0.0550 M $Ba(OH)_2$** solution. What is the concentration of the excess H^+ or OH^- ions left in this solution?

Answer: $Ba(OH)_2(aq) + 2\ HCl(aq) \rightarrow BaCl_2(aq) + 2\ H_2O(l)$; $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

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

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

$$= 1.24 \times 10^{-2}\ \text{mol Ba}^{2+} + 2.48 \times 10^{-2}\ \text{mol OH}^-$$

The net ionic equation requires a 1 : 1 mole ratio between OH^- and H^+. The actual mole OH^- to mole 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.

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

77. A student titrates an unknown amount of potassium hydrogen phthalate (**$KHC_8H_4O_4$**, often abbreviated KHP) with **20.46 mL** of a **0.1000-M NaOH** solution. KHP **(molar mass = 204.22 g/mol)** has one acidic hydrogen. What mass of KHP was titrated (reacted completely) by the sodium hydroxide solution?

Answer: KHP is a monoprotic acid: $NaOH(aq) + KHP(aq) \rightarrow H_2O(l) + NaKP(aq)$

$$\text{Mass KHP} = 0.02046\ \text{L NaOH} \times \frac{0.1000\ \text{mol NaOH}}{\text{L NaOH}} \times \frac{1\ \text{mol KHP}}{\text{mol NaOH}} \times \frac{204.22\ \text{g KHP}}{\text{mol KHP}}$$

$$= 0.4178\ \text{g KHP}$$

79. Assign oxidation states for all atoms in each of the following compounds.

a. $\mathbf{KMnO_4}$

b. $\mathbf{NiO_2}$

c. $\mathbf{Na_4Fe(OH)_6}$

d. $\mathbf{(NH_4)_2HPO_4}$

e. $\mathbf{P_4O_6}$

f. $\mathbf{Fe_3O_4}$

g. $\mathbf{XeOF_4}$

h. $\mathbf{SF_4}$

i. $\mathbf{CO}$

j. $\mathbf{C_6H_{12}O_6}$

Answer: Apply the rules in Table 4.2.

a. $KMnO_4$ is composed of K^+ and MnO_4^- ions. Assign oxygen an oxidation state 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. $Na_4Fe(OH)_6$ is composed of Na^+ cations and $Fe(OH)_6^{4-}$anions. $Fe(OH)_6^{4-}$ is composed of an iron cation and 6 OH^- anions. For an overall anion charge of 4−, iron must have a +2 oxidation state. As is usually the case in compounds, assign O a −2 oxidation state and H a +1 oxidation state. Na, +1; Fe, +2; O, −2; H, +1.

d. $(NH_4)_2HPO_4$ is made of NH_4^+ cations and HPO_4^{2-} anions. Assign +1 as the oxidation state of H and −2 as the oxidation state of O. In NH_4^+, $x + 4(+1) = +1$, $x = -3$ = oxidation state of N. In HPO_4^{2-}, $+1 + y + 4(-2) = -2$, $y = +5$ = oxidation state of P.

e. O, −2; P, +3

f. O, −2; $3x + 4(-2) = 0$, $x = +8/3$ = oxidation state of Fe; this is the average oxidation state of the three iron ions in Fe_3O_4. In the actual formula unit, there are two Fe^{3+} ions and one Fe^{2+} ion.

g. O, −2; F, −1; Xe, +6

h. F, −1; S, +4

i. O, −2; C, +2

j. H, +1; O, −2; C, 0

83. Specify which of the following are oxidation–reduction reactions, and identify the oxidizing agent, the reducing agent, the substance being oxidized, and the substance being reduced.

a. $Cu(s) + 2Ag^+(aq) \rightarrow 2Ag(s) + Cu^{2+}(aq)$

b. $HCl(g) + NH_3(g) \rightarrow NH_4Cl(s)$

c. $SiCl_4(l) + 2H_2O(l) \rightarrow 4HCl(aq) + SiO_2(s)$

d. $SiCl_4(l) + 2Mg(s) \rightarrow 2MgCl_2(s) + Si(s)$

e. $Al(OH)_4^-(aq) \rightarrow AlO_2^-(aq) + 2H_2O(l)$

Answer: 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 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	Ag^+	Cu	Cu	Ag^+
b.	No	–	–	–	–
c.	No	–	–	–	–
d.	Yes	$SiCl_4$	Mg	Mg	$SiCl_4$ (Si)
e.	No	–	–	–	–

In b, c, and e, no oxidation numbers change.

87. Balance each of the following oxidation–reduction reactions by using the oxidation states method.

a. $C_2H_6(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$

b. $Mg(s) + HCl(aq) \rightarrow Mg^{2+}(aq) + Cl^-(aq) + H_2(g)$

c. $Co^{3+}(aq) + Ni(s) \rightarrow Co^{2+}(aq) + Ni^{2+}(aq)$

d. $Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$

Answer:

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

$$\overset{-3\ +1}{C_2H_6} + \overset{0}{O_2} \rightarrow \overset{+4\ -2}{CO_2} + \overset{+1\ -2}{H_2O}$$

Each carbon atom changes from –3 to +4, an increase of 7. Each oxygen atom changes from 0 to –2, a decrease of 2. We need 7/2 O atoms for every C atom in order to balance electron gain with electron loss.

$$C_2H_6 + 7/2\ O_2 \rightarrow CO_2 + H_2O$$

Balancing the remainder of the equation by inspection:

$$C_2H_6(g) + 7/2\ O_2(g) \rightarrow 2\ CO_2(g) + 3\ H_2O(g)$$

or

$$2\ C_2H_6(g) + 7\ O_2(g) \rightarrow 4\ CO_2(g) + 6\ H_2O(g)$$

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 in order to balance the electrons transferred. The balanced equation is:

$$Mg(s) + 2\ HCl(aq) \rightarrow Mg^{2+}(aq) + 2\ Cl^-(aq) + H_2(g)$$

c. The oxidation state of nickel increases by 2 (0 to +2), and the oxidation state of cobalt decreases by 1 (+3 to +2). We need 2 Co^{3+} ions for every Ni atom in order to balance electron gain with electron loss. The balanced equation is:

$$Ni(s) + 2\ Co^{3+}(aq) \rightarrow Ni^{2+}(aq) + 2\ Co^{2+}(aq)$$

d. The equation is balanced (mass and charge balanced). Each hydrogen atom gains one electron (+1 → 0), and each zinc atom loses two electrons (0 → +2). We need 2 H atoms for every Zn atom in order to balance the electrons transferred. This is the ratio in the given equation:

$$Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$$