4.40 (a) $HC_2H_3O_2(aq) + KOH(aq) \rightarrow KC_2H_3O_2(aq) + H_2O(l)$

$HC_2H_3O_2(aq) + OH^-(aq) \rightarrow C_2H_3O_2^-(aq)\ H_2O(l)$

(b) $Cr(OH)_3(s) + 3HNO_3(aq) \rightarrow Cr(NO_3)_3(aq) + 3H_2O(l)$

$Cr(OH)_3(s) + 3H^+(aq) \rightarrow 3H_2O(l) + Cr^{3+}(aq)$

(c) $Ca(OH)_2(aq) + 2HClO(aq) \rightarrow Ca(ClO)_2(aq) + 2H_2O(l)$

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

4.41 *Analyze.* Given: names of reactants. Find: gaseous products.

Plan. Write correct chemical formulas for the reactants, complete and balance the metathesis reaction, and identify either H_2S or CO_2 products as gases. *Solve.*

(a) $CdS(s) + H_2SO_4(aq) \rightarrow CdSO_4(aq) + H_2S(g)$

$CdS(s) + 2H^+(aq) \rightarrow H_2S(g) + Cd^{2+}(aq)$

(b) $MgCO_3(s) + 2HClO_4(aq) \rightarrow Mg(ClO_4)_2(aq) + H_2O(l) + CO_2(g)$

$MgCO_3(s) + 2H^+(aq) \rightarrow H_2O(l) + CO_2(g) + Mg^{2+}(aq)$

4.42 (a) $FeO(s) + 2H^+(aq) \rightarrow H_2O(l) + Fe^{2+}(aq)$

(b) $NiO(s) + 2H^+(aq) \rightarrow H_2O(l) + Ni^{2+}(aq)$

4.43 *Analyze.* Given the formulas or names of reactants, write balanced molecular and net ionic equations for the reactions.

Plan. Write correct chemical formulas for all reactants. Predict products of the neutralization reactions by exchanging ion partners. Balance the complete molecular equation, identify spectator ions by recognizing strong electrolytes, write the corresponding net ionic equation (omitting spectators). *Solve.*

(a) $CaCO_3(s) + 2HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + H_2O(l) + CO_2(g)$

$2H^+(aq) + CaCO_3(s) \rightarrow H_2O(l) + CO_2(g) + Ca^{2+}(aq)$

(b) $FeS(s) + 2HBr(aq) \rightarrow FeBr_2(aq) + H_2S(g)$

$2H^+(aq) + FeS(s) \rightarrow H_2S(g) + Fe^{2+}(aq)$

4.44 $K_2O(aq) + H_2O(l) \rightarrow 2KOH(aq)$, molecular; $O^{2-}(aq) + H_2O(l) \rightarrow 2OH^-(aq)$, net ionic

base: (H^+ ion acceptor) $O^{2-}(aq)$; acid: (H^+ ion donor) $H_2O(aq)$; spectator: K^+

Oxidation-Reduction Reactions

4.45 (a) In terms of electron transfer, *oxidation* is the loss of electrons by a substance, and *reduction* is the gain of electrons (LEO says GER).

(b) Relative to oxidation numbers, when a substance is oxidized, its oxidation number increases. When a substance is reduced, its oxidation number decreases.

4.46 Oxidation and reduction can only occur together, not separately. When a metal reacts with oxygen, the metal atoms lose electrons and the oxygen atoms gain electrons. Free electrons do not exist under normal conditions. If electrons are lost by one substance they must be gained by another, and vice versa.

4.47 *Analyze.* Given the labeled periodic chart, determine which region is most readily oxidized and which is most readily reduced.

Plan. Review the definition of oxidation and apply it to the properties of elements in the indicated regions of the chart.

Solve. Oxidation is loss of electrons. Elements readily oxidized form positive ions; these are metals. Elements not readily oxidized tend to gain electrons and form negative ions; these are nonmetals. Elements in regions A, B, and C are metals, and their ease of oxidation is shown in Table 4.5. Metals in region A, Na, Mg, K, and Ca are most easily oxidized. Elements in region D are nonmetals and are least easily oxidized.

4.48 Elements (metals) from Table 4.5 in region A include Na, Mg, K, and Ca; those from region C are Hg and Au. Let's consider K and Au. Since metals from region A are more readily oxidized than those from region C, K will be oxidized to K^+ and Au^{3+} will be reduced to Au in the redox reaction. (Choose Au^{3+} because it is the Au ion shown in Table 4.5.)

In a balanced redox reaction, the number of electrons lost must equal the number of electrons gained. Since K loses 1 electron in forming K^+, while Au^{3+} gains 3 electrons when forming Au, 3 K atoms must be oxidized for every 1 Au^{3+} that is reduced. This relationship dictates the coefficients in the balanced redox reaction.

$$3K(s) + Au^{3+}(aq) \rightleftharpoons Au(s) + 3K^{+}(aq)$$

4.49 *Analyze.* Given the chemical formula of a substance, determine the oxidation number of a particular element in the substance.

Plan. Follow the logic in Sample Exercise 4.8. *Solve.*

(a) +4 (b) +4 (c) +7 (d) +1 (e) 0 (f) -1 (O_2^{2-} is peroxide ion)

4.50 (a) +4 (b) +2 (c) +3 (d) -2 (e) +3 (f) +6

4.51 *Analyze.* Given: chemical reaction. Find: element oxidized or reduced. *Plan.* Assign oxidation numbers to all species. The element whose oxidation number becomes more positive is oxidized; the one whose oxidation number decreases is reduced. *Solve.*

(a) $Ni \rightarrow Ni^{2+}$, Ni is oxidized; $Cl_2 \rightarrow 2Cl^-$, Cl is reduced

(b) $Fe^{2+} \rightarrow Fe$, Fe is reduced; $Al \rightarrow Al^{3+}$, Al is oxidized

(c) $Cl_2 \rightarrow 2Cl$, Cl is reduced; $2I^- \rightarrow I_2$, I is oxidized

(d) $S^{2-} \rightarrow SO_4^{2-}$(S, +6), S is oxidized; H_2O_2 (O, -1) $\rightarrow$ H_2O (O, -2); O is reduced

4.52 (a) acid-base reaction

(b) oxidation-reduction reaction; Fe is reduced, C is oxidized

(c) precipitation reaction

(d) oxidation-reduction reaction; Zn is oxidized, N is reduced

4.53 *Analyze.* Given: reactants. Find: balanced molecular and net ionic equations.

Plan. Metals oxidized by H^+ form cations. Predict products by exchanging cations and balance. The anions are the spectator ions and do not appear in the net ionic equations. *Solve.*

(a) $Mn(s) + H_2SO_4(aq) \rightarrow MnSO_4(aq) + H_2(g)$; $Mn(s) + 2H^+(aq) \rightarrow Mn^{2+}(aq) + H_2(g)$

Products with the metal in a higher oxidation state are possible, depending on reaction conditions and acid concentration.

(b) $2Cr(s) + 6HBr(aq) \rightarrow 2CrBr_3(aq) + 3H_2(g)$; $2Cr(s) + 6H^+(aq) \rightarrow 2Cr^{3+}(aq) + 3H_2(g)$

(c) $Sn(s) + 2HCl(aq) \rightarrow SnCl_2(aq) + H_2(g)$; $Sn(s) + 2H^+(aq) \rightarrow Sn^{2+}(aq) + H_2(g)$

(d) $2Al(s) + 6HCHO_2(aq) \rightarrow 2Al(CHO_2)_3(aq) + 3H_2(g)$;

$2Al(s) + 6HCHO_2(aq) \rightarrow 2Al^{3+}(aq) + 6CHO_2^-(aq) + 3H_2(g)$

4.54 (a) $2HCl(aq) + Ni(s) \rightarrow NiCl_2(aq) + H_2(g)$; $Ni(s) + 2H^+(aq) \rightarrow Ni^{2+}(aq) + H_2(g)$

(b) $H_2SO_4(aq) + Fe(s) \rightarrow FeSO_4(aq) + H_2(g)$; $Fe(s) + 2H^+(aq) \rightarrow Fe^{2+}(aq) + H_2(g)$

Products with the metal in a higher oxidation state are possible, depending on reaction conditions and acid concentration.

(c) $2HBr(aq) + Mg(s) \rightarrow MgBr_2(aq) + H_2(g)$; $Mg(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$

(d) $2HC_2H_3O_2(aq) + Zn(s) \rightarrow Zn(C_2H_3O_2)_2(aq) + H_2(g)$;

$Zn(s) + 2HC_2H_3O_2(aq) \rightarrow Zn^{2+}(aq) + 2C_2H_3O_2^-(aq) + H_2(g)$

4.55 *Analyze.* Given: a metal and an aqueous solution. Find: balanced equation.

Plan. Use Table 4.5. If the metal is above the aqueous solution, reaction will occur; if the aqueous solution is higher, NR. If reaction occurs, predict products by exchanging cations (a metal ion or H^+), then balance the equation. *Solve.*

(a) $Fe(s) + Cu(NO_3)_2(aq) \rightarrow Fe(NO_3)_2(aq) + Cu(s)$

(b) $Zn(s) + MgSO_4(aq) \rightarrow NR$

(c) $Sn(s) + 2HBr(aq) \rightarrow SnBr_2(aq) + H_2(g)$

(d) $H_2(g) + NiCl_2(aq) \rightarrow NR$

(e) $2Al(s) + 3CoSO_4(aq) \rightarrow Al_2(SO_4)_3(aq) + 3Co(s)$

4.56 (a) $Mn(s) + NiCl_2(aq) \rightarrow MnCl_2(aq) + Ni(s)$

(b) $Cu(s) + Cr(C_2H_3O_2)(aq) \rightarrow NR$

(c) $2Cr(s) + 3NiSO_4(aq) \rightarrow Cr_2(SO_4)_3(aq) + 3Ni(s)$

(d) $Pt(s) + HBr(aq) \rightarrow NR$

(e) $H_2(g) + CuCl_2(aq) \rightarrow Cu(s) + 2HCl(aq)$

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

ii. $Cd(s) + Ni^{2+}(aq) \rightarrow Ni(s) + Cd^{2+}(aq)$

(b) According to Table 4.5, the most active metals are most easily oxidized, and Zn is more active than Ni. 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.

(c) Place an iron strip in $CdCl_2(aq)$. 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 (a) $Br_2 + 2NaI \rightarrow 2NaBr + I_2$ indicates that Br_2 is more easily reduced than I_2.

$Cl_2 + 2NaBr \rightarrow 2NaCl + Br_2$ shows that Cl_2 is more easily reduced than Br_2.

The order for ease of reduction is $Cl_2 > Br_2 > I_2$. Conversely, the order for ease of oxidation is $I^- > Br^- > Cl^-$.

(b) Since the halogens are nonmetals, they tend to form anions when they react chemically. Nonmetallic character decreases going down a family and so does the tendency to gain electrons during a chemical reaction. Thus, the ease of reduction of the halogen, X_2, decreases going down the family and the ease of oxidation of the halide, X^-, increases going down the family.

(c) $Cl_2 + 2KI \rightarrow 2KCl + I_2$; $Br_2 + LiCl \rightarrow$ no reaction

Solution Composition; Molarity

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 (a) The concentration of the remaining solution is unchanged, assuming the original solution was thoroughly mixed. Molar concentration is a ratio of moles solute to liters solution. Although there are fewer moles solute remaining in the flask, there is also less solution volume, so the ratio of moles solute/solution volume remains the same.

(b) The second solution is five times as concentrated as the first. An equal volume of the more concentrated solution will contain five times as much solute (five times the number of moles and also five times the mass) as the 0.50 *M* solution. Thus, the mass of solute in the 2.50 *M* solution is 5×4.5 g = 22.5 g.

Mathematically: