21. During a titration, 0.200*M* HCl is added to a NaOH solution of unknown concentration. What is the concentration of the NaOH solution if 20.0 mL of it is neutralized by 30.7 mL of the standard solution?

$NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$;
0.0307 L HCl × 0.200 mol HCl/L HCl
= 6.14×10^{-3} mol HCl;
mol HCl : mol NaOH = 1:1,
so 6.14×10^{-3} mol HCl : 6.14×10^{-3} mol NaOH;
6.14×10^{-3} mol NaOH/0.0200 L NaOH = 0.307*M*

22. A 25.0-mL sample of H_2SO_4 is neutralized by 27.4 mL of 1.00*M* KOH. What is the concentration of the acid?

$2KOH(aq) + H_2SO_4(aq) \rightarrow K_2SO_4(aq) + 2H_2O(l)$;
0.0274 L KOH × 1.00 mol KOH/L KOH
= 0.0274 mol KOH; mol KOH : mol H_2SO_4 = 2:1,
so 0.0274 mol KOH : 0.0137 mol H_2SO_4;
0.0137 mol H_2SO_4/0.0250 L H_2SO_4 = 0.548*M*

23. A 50.0-mL sample of 0.0100*M* $Ca(OH)_2$ is neutralized by 45.6 mL of HBr. What is the molarity of the acid?

$Ca(OH)_2(aq) + 2HBr(aq) \rightarrow CaBr_2(aq) + 2H_2O(l)$;
0.0500 L $Ca(OH)_2$ × 0.0100 mol $Ca(OH)_2$/L $Ca(OH)_2$
= 5.00×10^{-4} mol $Ca(OH)_2$;
mol $Ca(OH)_2$: mol HBr = 1:2,
so 5.00×10^{-4} mol $Ca(OH)_2$: 1.00×10^{-3} mol HBr;
1.00×10^{-3} mol HBr/0.0456 L HBr = 0.0219*M*

Chapter 20

Determine the oxidation number of the boldface element in these ions.

1. $\mathbf{HgCl_4^-}$

+3

2. $\mathbf{NO_2}$

+4

3. $\mathbf{MnO_2}$

+4

4. metallic **Au**

0

5. $\mathbf{Na_2SiF_6}$

+4

6. $\mathbf{Zn(NO_3)_2}$

+5

7. $\mathbf{Mg_3P_2}$

−3

8. $\mathbf{Na_3PO_4}$

+5

9. $\mathbf{H_2O_2}$

−1

10. $\mathbf{ClO_3^-}$

+5

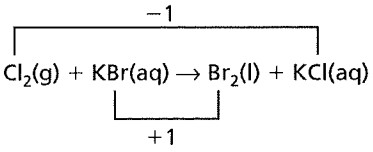
Balance the following equations, using the oxidation number method for the redox part of the equation. Show your work.

11. $Cu_2O(s) + H_2(g) \rightarrow Cu(s) + H_2O(l)$

$$\overset{\overset{2(-1) = -2}{\frown}}{Cu_2O(s) + H_2(g) \rightarrow 2Cu(s) + H_2O(l)}$$
$$\underset{\underset{2(+1) = +2}{\smile}}{}$$

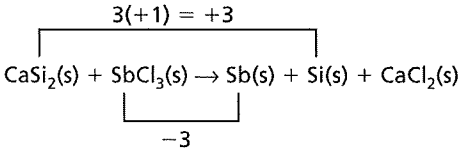
$Cu_2O(s) + H_2(g) \rightarrow 2Cu(s) + H_2O(l)$

12. $Cl_2(g) + KBr(aq) \longrightarrow Br_2(l) + KCl(aq)$



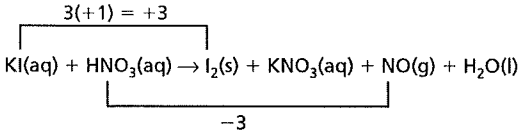
$Cl_2(g) + 2KBr(aq) \longrightarrow Br_2(l) + 2KCl(aq)$

13. $CaSi_2(s) + SbCl_3(s) \longrightarrow Sb(s) + Si(s) + CaCl_2(s)$



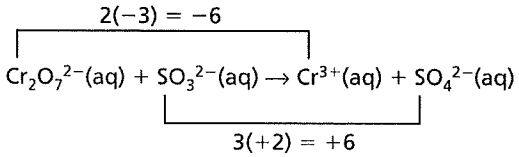
$3CaSi_2(s) + 2SbCl_3(s) \longrightarrow 2Sb(s) + 6Si(s) + 3CaCl_2(s)$

14. $KI(aq) + HNO_3(aq) \longrightarrow I_2(s) + KNO_3(aq) + NO(g) + H_2O(l)$



$6KI(aq) + 8HNO_3(aq) \longrightarrow 3I_2(s) + 6KNO_3(aq) + 2NO(g) + 4H_2O(l)$

15. $Cr_2O_7^{2-}(aq) + SO_3^{2-}(aq) \longrightarrow Cr^{3+}(aq) + SO_4^{2-}(aq)$ in an acidic solution



$Cr_2O_7^{2-}(aq) + 3SO_3^{2-}(aq) + 8H^+(aq) \longrightarrow 2Cr^{3+}(aq) + 3SO_4^{2-}(aq) + 4H_2O(l)$

Write half-reactions for each of the following redox reactions. Identify each half-reaction as being either oxidation or reduction.

16. $SnS_2(s) + O_2(g) \longrightarrow SnO_2(s) + SO_2(g)$

$S^{2-} \longrightarrow S^{4+} + 6e^-$, oxidation;
$O_2 + 4e^- \longrightarrow 2O^{2-}$, reduction

17. $Mg(s) + N_2(g) \longrightarrow Mg_3N_2(s)$

$Mg \longrightarrow Mg^{2+} + 2e^-$, oxidation;
$N_2 + 6e^- \longrightarrow 2N^{3-}$, reduction

18. $Al(s) + Cl_2(g) \longrightarrow AlCl_3(s)$

$Al \longrightarrow Al^{3+} + 3e^-$, oxidation;
$Cl_2 + 2e^- \longrightarrow 2Cl^-$, reduction

19. $NH_3(aq) + PbO(s) \longrightarrow N_2(g) + Pb(s) + H_2O(l)$

$2N^{3-} \longrightarrow N_2 + 6e^-$, oxidation;
$Pb^{2+} + 2e^- \longrightarrow 2Pb$, reduction

20. $Cu_2S(s) + O_2(g) \longrightarrow Cu^{2+}(aq) + SO_4^{2-}(aq)$ (Hint: Two different elements are oxidized.)

$S^{2-} \longrightarrow S^{4+} + 6e^-$, $Cu^+ \longrightarrow Cu^{2+} + 1e^-$, oxidation;
$O_2 + 4e^- \longrightarrow 2O^{2-}$, reduction

Use your answers for questions 16–20 to help you balance the following equations, using half-reactions for the redox part of the equation. Show your work.

21. $SnS_2(s) + O_2(g) \longrightarrow SnO_2(s) + SO_2(g)$

$2(S^{2-} + S^{4+} + 6e^-)$; $3(O_2 + 4e^- \longrightarrow 2O^{2-})$

$2S^{2-} \longrightarrow + 2S^{4+} + 12e^-$
$3O_2 + 12e^- \longrightarrow 6O^{2-}$

$SnS_2(s) + 3O_2(g) \longrightarrow SnO_2(s) + 2SO_2(g)$

22. $Mg(s) + N_2(g) \longrightarrow Mg_3N_2(s)$

$3(Mg \longrightarrow Mg^{2+} + 2e^-)$; $N_2 + 6e^- \longrightarrow 2N^{3-}$

$3Mg \longrightarrow 3Mg^{2+} + 6e^-$

$N_2 + 6e^- \longrightarrow 2N^{3-}$

$3Mg(s) + N_2(g) \longrightarrow Mg_3N_2(s)$

23. $Al(s) + Cl_2(g) \longrightarrow AlCl_3(s)$

$2(Al \longrightarrow Al^{3+} + 3e^-)$; $3(Cl_2 + 2e^- \longrightarrow 2Cl^-)$

$2Al \longrightarrow 2Al^{3+} + 6e^-$

$3Cl_2 + 6e^- \longrightarrow 6Cl^-)$

$2Al(s) + 3Cl_2(g)\ 2AlCl_3(s)$

24. $NH_3(aq) + PbO(s) \longrightarrow N_2(g) + Pb(s) + H_2O(l)$

$2N^{3-} \longrightarrow N_2 + 6e^-$; $3(Pb^{2+} + 2e^- \longrightarrow 2Pb)$

$2N^{3-} \longrightarrow N_2 + 6e^-$

$3Pb^{2+} + 6e^- \longrightarrow 6Pb$

$2NH_3(aq) + 3PbO(s) \longrightarrow N_2(g) + 3Pb(s) + 3H_2O(l)$

25. $Cu_2S(s) + O_2(g) \longrightarrow Cu^{2+}(aq) + SO_4^{2-}(aq)$ in an acidic solution (Hint: Look at the ratio of the two oxidized elements in the equation.)

Cu^+ and S^{2-} are both oxidized. According to the equation, two Cu^+ ions are oxidized for every one S^{2-} ion oxidized, for a total loss of $10e^-$ for the oxidation part:

$S^{2-} \longrightarrow S^{6+} + 8e^-$

$2Cu^+ \longrightarrow 2Cu^{2+} + 2e^-$

reduction: $O_2 + 4e^- \longrightarrow 2O^{2-}$

Because 20 is the least common multiple of 10 and 4, multiply the oxidation equations by 2 and the reduction equation by 5.

$2S^{2-} \longrightarrow 2S^{6+} + 16e^-$

$4Cu^+ \longrightarrow 4Cu^{2+} + 4e^-$

$5O_2 + 20e^- \longrightarrow 10O^{2-}$

$2Cu_2S(s) + 5O_2(g) + 4H^+(aq) \longrightarrow$
$4Cu^{2+}(aq) + 2SO_4^{2-}(aq) + 2H_2O(l)$

Chapter 21

Use data from Table 21-1 as needed in the following problems. Assume that all half-cells are under standard conditions.

1. For each of these pairs of half-reactions, write a balanced equation for the overall cell reaction and calculate the standard cell potential, E^0_{cell}.

a. $Cs^+(aq) + e^- \longrightarrow Cs(s)$

$Cu^+(aq) + e^- \longrightarrow Cu(s)$

Cell reaction:

$Cu^+(aq) + Cs(s) \longrightarrow Cu(s) + Cs^+(aq)$

$E^0_{cell} =$

$+0.521\ V - (-3.026\ V) = +3.547\ V$

b. $Hg^{2+}(aq) + 2e^- \longrightarrow Hg(l)$

$Mn^{2+}(aq) + 2e^- \longrightarrow Mn(s)$

Cell reaction:

$Hg^{2+}(aq) + Mn(s) \longrightarrow Hg(l) + Mn^{2+}(aq)$

$E^0_{cell} =$

$+0.851\ V - (-1.185\ V) = +2.036\ V$

c. $Fe^{3+}(aq) + 3e^- \longrightarrow Fe(s)$

$Cr^{3+}(aq) + 3e^- \longrightarrow Cr(s)$

Cell reaction:

$Fe^{3+}(aq) + Cr(s) \longrightarrow Fe(s) + Cr^{3+}(aq)$

$E^0_{cell} =$

$-0.037\ V - (-0.744\ V) = +0.707\ V$

d. $Br_2(g) + 2e^- \longrightarrow 2Br^-(aq)$

$Au^+(aq) + e^- \longrightarrow Au(s)$

Cell reaction:

$2Au^+(aq) + 2Br^-(aq) \longrightarrow 2Au(s) + Br_2(g)$