First-Year Chemistry

2. Stoichiometry and Reactions

Solutions to further problems:

1. a) $CaCO_3$ = solid, HCl = aqueous, $CaCl_2$ = aqueous, CO_2 = gas, and H_2O = liquid

   b) $CaCO_{3\,(s)} + 2\,HCl_{(aq)} \rightarrow H_2O_{\,(g)} + CO_{2\,(g)}$

   c) 0.500 mol $CaCO_3$ $\left(\dfrac{1\,\text{mol}\,CO_2}{1\,\text{mol}\,CaCO_3}\right)$ = 0.500 mol of CO_2 [3 sig figs]

   d) 0.250 mol $CaCO_3$ $\left(\dfrac{1\,\text{mol}\,H_2O}{1\,\text{mol}\,CaCO_3}\right)\left(\dfrac{18.02\,\text{g}\,H_2O}{1\,\text{mol}\,H_2O}\right)$ = 4.50 g of H_2O [3 sig figs]

   e) 10.5 g $CaCO_3$ $\left(\dfrac{1\,\text{mol}\,CaCO_3}{100.09\,\text{g}\,CaCO_3}\right)\left(\dfrac{1\,\text{mol}\,CO_2}{1\,\text{mol}\,CaCO_3}\right)\left(\dfrac{44.02\,\text{g}\,CO_2}{1\,\text{mol}\,CO_2}\right)$ = 4.62 g of CO_2 [3 sig figs]

   f) 20.0 g $CaCO_3$ $\left(\dfrac{1\,\text{mol}\,CaCO_3}{100.09\,\text{g}\,CaCO_3}\right)\left(\dfrac{1\,\text{mol}\,H_2O}{1\,\text{mol}\,CaCO_3}\right)\left(\dfrac{6.02\,\text{X}\,10^{23}\ \text{molecules}\,H_2O}{1\,\text{mol}\,H_2O}\right)$ = 1.20 X 10^{23} molecules of water [3 sig figs]

   g) 14.0 g CO_2 $\left(\dfrac{1\,\text{mol}\,CO_2}{44.01\,\text{g}\,CO_2}\right)\left(\dfrac{2\,\text{mol}\,HCl}{1\,\text{mol}\,CO_2}\right)$ = 0.636 mol HCl [3 sig figs]

2. (13.7 cm^3 H_2SO_4)$\left(\dfrac{1.20\ \text{mol}\ H_2SO_4}{1000\ \text{cm}^3\ H_2SO_4}\right)\left(\dfrac{1}{0.100\ \text{dm}^3}\right)$ = 0.1644 = 0.164 M (3 sig figs)

3. (a) $[NH_4^+]$ = $\left(\dfrac{1.20\ \text{mol}\ (NH_4)_2SO_4}{1\ \text{dm}^3}\right)\left(\dfrac{2\ \text{mol}\ NH_4^+}{1\ \text{mol}\ (NH_4)_2SO_4}\right)$ = 2.40 M (3 sig figs)

   $[SO_4^{2-}]$ = $\left(\dfrac{1.20\ \text{mol}\ (NH_4)_2SO_4}{1\ \text{dm}^3}\right)\left(\dfrac{1\ \text{mol}\ SO_4^{2-}}{1\ \text{mol}\ (NH_4)_2SO_4}\right)$ = 1.20 M (3 sig figs)

   (b) $[Ca^{2+}]$ = $\left(\dfrac{0.250\ \text{mol}\ CaCl_2}{1\ \text{dm}^3}\right)\left(\dfrac{1\ \text{mol}\ Ca^{2+}}{1\ \text{mol}\ CaCl_2}\right)$ = 0.250 M (3 sig figs)

   $[Cl^-]$ = $\left(\dfrac{0.250\ \text{mol}\ CaCl_2}{1\ \text{dm}^3}\right)\left(\dfrac{2\ \text{mol}\ Cl^-}{1\ \text{mol}\ CaCl_2}\right)$ = 0.500 M (3 sig figs)

   (c) $[Fe^{2+}]$ = $\left(\dfrac{0.400\ \text{mol}\ Fe_2(SO_4)_3}{1\ \text{dm}^3}\right)\left(\dfrac{2\ \text{mol}\ Fe^{2+}}{1\ \text{mol}\ Fe_2(SO_4)_3}\right)$ = 0.800 M (3 sig figs)

   $[SO_4^{2-}]$ = $\left(\dfrac{0.400\ \text{mol}\ Fe_2(SO_4)_3}{1\ \text{dm}^3}\right)\left(\dfrac{3\ \text{mol}\ SO_4^{2-}}{1\ \text{mol}\ Fe_2(SO_4)_3}\right)$ = 1.20 M (3 sig figs)

4. The equation for the reaction is $2HClO_{3\,(aq)} + Ba(OH)_{2\,(aq)} \rightarrow Ba(ClO_3)_{2\,(aq)} + 2H_2O_{(l)}$
   Answer is 0.100 (same as the number of moles of $HClO_3$ since firstly it's the limiting reagent, and secondly, the ratio of $HClO_3$ to H_2O are equal).

5. Equation is: $AgNO_{3\,(aq)} + KBr_{(aq)} \rightarrow AgBr_{\,(s)} + KNO_{3\,(aq)}$

   (2.00 g $AgNO_3$)$\left(\dfrac{1\,\text{mol}\,AgNO_3}{169.88\ \text{g}\ AgNO_3}\right)\left(\dfrac{1\ \text{mol}\ AgBr}{1\,\text{mol}\,AgNO_3}\right)\left(\dfrac{187.77\ \text{g}\ AgBr}{1\ \text{mol}\ AgBr}\right)$ = 2.210 g = 2.21 g AgBr

   (4.00 g KBr)$\left(\dfrac{1\,\text{mol}\,KBr}{119.00\ \text{g}\ KBr}\right)\left(\dfrac{1\ \text{mol}\ AgBr}{1\,\text{mol}\,KBr}\right)\left(\dfrac{187.77\ \text{g}\ AgBr}{1\ \text{mol}\ AgBr}\right)$ = 6.311 g = 6.31 g AgBr

$AgNO_3$ is the limiting reagent. And therefore, 2.21 g of solid AgBr will be produced.

6. a) Yes, a reaction does take place (iron(III) hydroxide is insoluble)

   $Fe^{+3}{}_{(aq)} + 3OH^{-}{}_{(aq)} \rightarrow Fe(OH)_{3\,(s)}$

   b) Yes, a reaction does take place (silver carbonate is insoluble)

   $2Ag^{+}{}_{(aq)} + CO_3^{2-}{}_{(aq)} \rightarrow Ag_2CO_{3\,(s)}$

   c) No, a reaction does not take place; all acetates and nitrates are soluble.

   d) Yes.

   $MgCO_{3\,(s)} + 2H^{+}{}_{(aq)} \rightarrow Mg^{2+}{}_{(aq)} + CO_{2\,(g)} + H_2O_{\,(l)}$

   e) Yes.

   $NaHCO_{3\,(s)} + 2H^{+}{}_{(aq)} \rightarrow Na^{+}{}_{(aq)} + CO_{2\,(g)} + H_2O_{\,(l)}$

   f) Yes.

   $NaOH_{\,(s)} + CH_3COOH_{\,(aq)} \rightarrow Na^{+}{}_{(aq)} + CH_3COO^{-}{}_{(aq)} + H_2O_{\,(l)}$

   g) Yes.

   $NH_{3\,(aq)} + HBr_{\,(aq)} \rightarrow NH_4^{+}{}_{(aq)} + Br^{-}{}_{(aq)}$

   h) Yes.

   $MgO_{\,(s)} + 2H^{+}{}_{(aq)} \rightarrow Mg^{2+}{}_{(aq)} + H_2O_{\,(l)}$

7. a) Total charge on ion = –1

   $\Rightarrow$ charge on I + 3 (charge on O) = –1

   $\Rightarrow$ charge on I + 3 (–2) = –1

   $\Rightarrow$ charge on I – 6 = –1

   $\Rightarrow$ charge on I = –1 + 6

   $\Rightarrow$ charge on I = 5

   Oxidation number of I in iodate ion = 5+.

   b) Total charge on compound = 0

   $\Rightarrow$ charge on H + charge on Cl + charge on O = 0

   $\Rightarrow$ 1 + charge on Cl + (–2) = 0

   $\Rightarrow$ 1 + charge on Cl – 2 = 0

   $\Rightarrow$ charge on Cl = 2 – 1

   $\Rightarrow$ charge on Cl = 1

   Oxidation number of Cl in HClO = 1+

   c) Oxidation number of N in NO = 2+

   d) Oxidation number of N in HNO_3 = 5+

   e) Oxidation number of Mn in MnO_4^{-} = 7+

   f) Oxidation number of S in $S_2O_3^{2-}$ = 2+

g) Oxidation number of S in SO_4^{2-} = 6+

h) Oxidation number of Mn in MnO_4^{-} = 6+

8. a)

$$\begin{array}{lllllll}
4+ & \xrightarrow{\text{Reduction}} & & 0 & & & \\
\downarrow & & & \downarrow & & & \\
TiCl_4 & + & Mg & \rightarrow & Ti & + & MgCl_2 \\
 & & \uparrow & & & & \uparrow \\
 & & 0 & \xrightarrow[\text{Oxidation}]{} & & & 2+
\end{array}$$

Balancing each half-equation:

Oxidation half-equation: $Mg \rightarrow Mg^{2+} + 2e^{-}$

Reduction half-equation: $Ti^{4+} + 4e^{-} \rightarrow Ti$

Multiplying the oxidation half equation by 2 and adding the two balanced half-equations:

$$\begin{array}{rcl}
2Mg & \rightarrow & 2Mg^{2+} + 4e^{-} \\
Ti^{4+} + 4e^{-} & \rightarrow & Ti \\
\hline
Ti^{4+} + 2Mg & \rightarrow & 2Mg^{2+} + Ti \\
\hline
\end{array}$$

Adding and balancing the rest of the elements:

$TiCl_{4\,(g)} + 2\,Mg_{(l)} \rightarrow 2\,MgCl_{2(s)} + Ti_{(s)}$

b)

$$\begin{array}{lllllll}
0 & \xrightarrow{\text{Oxidation}} & & 1- & & & \\
\downarrow & & & \downarrow & & & \\
Cl_2 & + & Br^{-} & \rightarrow & Cl^{-} & + & Br_2 \\
 & & \uparrow & & & & \uparrow \\
 & & 1- & \xrightarrow[\text{Reduction}]{} & & & 0
\end{array}$$

Balancing each half-equation:

Oxidation half-equation: $2Br^{-} \rightarrow Br_2 + 2e^{-}$

Reduction half-equation: $Cl_2 + 2e^{-} \rightarrow Cl^{-}$

Multiplying the oxidation half equation by 2 and adding the two balanced half-equations:

$$\begin{array}{rcl}
4Br^{-} & \rightarrow & Br_2 + 2e^{-} \\
Cl_2 + 2e^{-} & \rightarrow & 2Cl^{-} \\
\hline
Cl_{2\,(g)} + 2Br^{-}_{(aq)} & \rightarrow & 2Cl^{-}_{(aq)} + Br_{2\,(l)} \\
\hline
\end{array}$$

c)

$$\begin{array}{lllllll}
3- & \xrightarrow{\text{Oxidation}} & & 2+ & & & \\
\downarrow & & & \downarrow & & & \\
NH_3 & + & O_2 & \rightarrow & NO & + & H_2O \\
 & & \uparrow & & \uparrow & & \uparrow \\
 & & 0 & \xrightarrow[\text{Reduction}]{} & 2- & & 2-
\end{array}$$

Balancing each half-equation:

Oxidation half-equation: $N^{3-} \rightarrow N^{2+} + 5e^{-}$

Reduction half-equation: $O_2 + 4e^{-} \rightarrow 2O^{2-}$

Multiplying the oxidation half-equation by 4 and the reduction half-equation by 5, and adding the two balanced half-equations:

$$\begin{array}{rcl}
4N^{3-} & \rightarrow & 4N^{2+} + 20e^{-} \\
5O_2 + 20e^{-} & \rightarrow & 10O^{2-} \\
\hline
4N^{3-} + 5O_2 & \rightarrow & 4N^{2+} + 10O^{2-} \\
\hline
\end{array}$$

Adding and balancing the rest of the elements:

$4NH_{3(g)} + 5O_{2\ (aq)} \rightarrow 4NO_{\ (g)} + 6H_2O_{\ (l)}$

d)

$$\begin{array}{lllllllll}
2+ & \xrightarrow{\text{Reduction}} & & 0 & & & & & \\
\downarrow & & & \downarrow & & & & & \\
CuO & + & NH_3 & \rightarrow & Cu & + & N_2 & + & H_2O \\
 & & \uparrow & & & & \uparrow & & \\
 & & 3- & \xrightarrow[\text{Oxidation}]{} & & & 0 & &
\end{array}$$

Balancing each half-equation:

Oxidation half-equation: $2N^{3-} \rightarrow N_2 + 6e^{-}$

Reduction half-equation: $Cu^{2+} + 2e^{-} \rightarrow Cu$

Multiplying the reduction half-equation by 3 and adding the two balanced half-equations:

$$\begin{array}{rcl}
2N^{3-} & \rightarrow & N_2 + 6e^{-} \\
3Cu^{2+} + 6e^{-} & \rightarrow & 3Cu \\
\hline
3Cu^{2+} + 2N^{3-} & \rightarrow & N_2 + 3Cu \\
\hline
\end{array}$$

Adding and balancing the rest of the elements

$3CuO_{\ (s)} + 2NH_{3(g)} \rightarrow 3Cu_{\ (s)} + N_{2\ (g)} + 3H_2O_{\ (l)}$

e)

$$\begin{array}{ccccc} 2- & \xrightarrow{\text{Reduction}} & 3- & & \\ \downarrow & & \downarrow & & \\ P_2H_4 & \rightarrow & PH_3 & + & P_4H_2 \\ \uparrow & & & & \uparrow \\ 2- & & \xrightarrow[\text{Oxidation}]{} & & -\frac{1}{2} \end{array}$$

Balancing each half-equation:

Oxidation half-equation: $2P_2^{2-} \rightarrow P_4^{-\frac{1}{2}} + 6e^-$

Reduction half-equation: $3P_2^{2-} + 6e^- \rightarrow 6P^{3-}$

Adding the two balanced half-equations:

$$\begin{array}{rcl} 2P_2^{2-} & \rightarrow & P_4^{-\frac{1}{2}} + 6e^- \\ 2P_2^{2-} + 6e^- & \rightarrow & 6P^{3-} \\ \hline 5P_2^{2-} & \rightarrow & 6P^{3-} + P_4^{-\frac{1}{2}} \\ \hline \end{array}$$

Adding and balancing the rest of the elements:

$5P_2H_{4\,(g)} \rightarrow 6PH_{3\,(g)} + P_4H_{2\,(g)}$

9. The first two solutions are show all the details.

a) Half-equation method: First, the skeletal equation:

$$\begin{array}{ccccccc} 7+ & & \xrightarrow{\text{Reduction}} & & 2+ & & \\ \downarrow & & & & \downarrow & & \\ MnO_4^-{}_{(aq)} & + & H_2C_2O_{4\,(aq)} & \rightarrow & Mn^{2+}{}_{(aq)} & + & CO_{2\,(g)} \\ & & \uparrow & & & & \uparrow \\ & & 3+ & \xrightarrow{\text{Oxidation}} & & & 4+ \end{array}$$

Oxidation number of manganese decreases from 7+ to 2+, hence MnO_4^{2-} is the oxidizing agent.

Oxidation number of carbon increases from 3+ to 4+ hence, oxalic acid is the reducing agent.

1. To start balancing the equation, the two skeletal equations for the half reactions.

Reduction half-reaction: $MnO_4^- \rightarrow Mn^{2+}$

Oxidation half-reduction: $H_2C_2O_4 \rightarrow CO_2$

2. Balance by inspection all the elements in the half reactions except O, H, and the charge.

Reduction half-reaction: $MnO_4^- \rightarrow Mn^{2+}$

Oxidation half-reduction: $H_2C_2O_4 \rightarrow 2\ CO_2$

3. Balance the oxygen and hydrogen atoms. For oxanions in *acidic* solution, we balance H and O by adding H^+ and H_2O:

i) First balance the O atoms by adding H_2O:

Reduction half-reaction. Add 4 H_2O on the right to balance the four O atoms on the left:

$$MnO_4^- \rightarrow Mn^{2+} + 4\ H_2O$$

Oxidation half reaction: The O atoms are already balanced.

$$H_2C_2O_4 \rightarrow 2\ CO_2$$

ii) Next, add H^+ to balance the H atoms on either side of the half reactions:

Reduction half-reaction: Add 8 H^+ on the left to balance the eight H atoms on the right:

$$MnO_4^- + 8\ H^+ \rightarrow Mn^{2+} + 4\ H_2O$$

Oxidation half reaction. The equation needs 2 H^+ on the right to balance the H atoms in the $H_2C_2O_2$ molecule on the left.

$$H_2C_2O_4 \rightarrow 2\ CO_2 + 2\ H^+$$

4. Balance the electric charges by adding electrons.

Reduction half-reaction. To balance the 2+ charge on the right and 7+ charge on the left, we need to add five electrons to the left:

$$MnO_4^- + 8\ H^+ + 5\ e^- \rightarrow Mn^{2+} + 4\ H_2O$$

Oxidation half reaction. We need to add two electrons on the right to cancel the two positive charges.

$$H_2C_2O_4 \rightarrow 2\ CO_2 + 2\ H^+ + 2\ e^-$$

5. prepare the two half reactions for summation by making the number of electrons the same in both. (Electrons cannot be created or destroyed.) Multiply the reduction half-reaction by 2 and the oxidation half-reaction by 5:

Reduction half-reaction: $2\ MnO_4^- + 16\ H^+ + 10\ e^- \rightarrow 2\ Mn^{2+} + 8\ H_2O$

Oxidation half reaction: $5\ H_2C_2O_4 \rightarrow 10\ CO_2 + 10\ H^+ + 10\ e^-$

6. Combine the two half-reactions.

i) First, add the two half-reactions in step 5:

$$2\ MnO_4^- + 16\ H^+ + 10\ e^- + 5\ H_2C_2O_4 \rightarrow 2\ Mn^{2+} + 8\ H_2O + 10\ CO_2 + 10\ H^+ + 10\ e^-$$

ii) Next, simplify the balanced equation by canceling like species on both sides of the equation. In this example, $10e^-$ and 10 H^+ can be canceled.

$$2\ MnO_4^- + 6\ H^+ + 5\ H_2C_2O_4 \rightarrow 2\ Mn^{2+} + 8\ H_2O + 10\ CO_2$$

iii) Finally, insert the state of each species:

$$2\ MnO_4^-{}_{(aq)} + 6\ H^+{}_{(aq)} + 5\ H_2C_2O_4{}_{(aq)} \rightarrow 2\ Mn^{2+}{}_{(aq)} + 8\ H_2O_{(l)} + 10\ CO_2{}_{(g)}$$

The expression is the fully balanced net ionic equation.

a) Oxidation state method: First, the skeletal equation:

$$\begin{array}{ccccccc} 7+ & & \xrightarrow{\text{Reduction}} & & 2+ & & \\ \downarrow & & & & \downarrow & & \\ MnO_4^-{}_{(aq)} & + & H_2C_2O_4{}_{(aq)} & \rightarrow & Mn^{2+}{}_{(aq)} & + & CO_2{}_{(g)} \\ & & \uparrow & & & & \uparrow \\ & & 3+ & \xrightarrow{\text{Oxidation}} & & & 4+ \end{array}$$

Step 1: Balancing the total change in oxidation numbers

Total change in oxidation number of manganese = final – initial = 2 – 7 = – 5

Total change in oxidation number of carbon = 2(4) – 2(3) = 2

Total change in oxidation number of Mn + total change in oxidation number of C = 0

$(-5)x + (2)y = 0$

Therefore, $x = 2$, $y = 5$

$$2\ MnO_4^-{}_{(aq)} + 5\ H_2C_2O_4{}_{(aq)} \rightarrow 2\ Mn^{2+}{}_{(aq)} + 10\ CO_2{}_{(g)}$$

Step 2: Balancing charges by adding H^+.

Charge on left hand side = –2

Charge on right hand side = +4

Since we are balancing charges by adding a positively charged species, the left hand side

charge must be brought up to +4 which requires $6H^+$.

$$2\ MnO_4^-{}_{(aq)} + 6\ H^+{}_{(aq)} + 5\ H_2C_2O_4{}_{(aq)} \rightarrow 2\ Mn^{2+}{}_{(aq)} + 10\ CO_2{}_{(g)}$$

Step 3: Balancing excess hydrogen and oxygen by adding H_2O.

Number of H atoms on the left hand side = 6 + 10 = 16
Number of H atoms on the right hand side = 0

Right hand requires 16 H atoms.

Number of O atoms on the left hand side = 2 (4) + 5 (4) = 28
Number of O atms on the right hand side = 10 (2) = 20

Right hand requires 8 O atoms.

Therefore, right hand side requires $8H_2O$ (16 H and 8 O atoms)

$$2\ MnO_4^-{}_{(aq)} + 6\ H^+{}_{(aq)} + 5\ H_2C_2O_4{}_{(aq)} \rightarrow 2\ Mn^{2+}{}_{(aq)} + 8\ H_2O_{(l)} + 10\ CO_2{}_{(g)}$$

b) The skeletal equation:

$$\begin{array}{ccccccc} 7+ & \xrightarrow{\text{Reduction}} & & & 4+ & & \\ \downarrow & & & & \downarrow & & \\ MnO_4^-{}_{(aq)} & + & Br^-{}_{(aq)} & \rightarrow & MnO_2{}_{(s)} & + & BrO_3^-{}_{(aq)} \\ & & \uparrow & & & & \uparrow \\ & & 1- & \xrightarrow{\text{Oxidation}} & & & 5+ \end{array}$$

Since the oxidation number of manganese decreases from 7+ to 4+ it is the oxidizing agent.

The oxidation number of bromine increases from 1– to 5+ it is the reducing agent.

The procedure is similar to that balancing redox equations in acidic solution, but step 3 is a little more involved.

1. Write the skeletal equations for the half-reactions.

Reduction half-reaction: $MnO_4^- \rightarrow MnO_2$

Oxidation half-reduction: $Br^- \rightarrow BrO_3^-$

2. Balance by inspection all elements in the half-reactions except O, H and the charge. The equations of both half-reactions are already balanced in this case.

3. Balance the oxygen and hydrogen atoms. For oxoanions in *basic* solution, we balance H and O by adding OH^- and H_2O.

i) First balance the O atoms by adding H_2O:

Reduction half-reaction. Add 2 H_2O on the right side to balance the two extra O atoms on the left.

$$MnO_4^- \rightarrow MnO_2 + 2H_2O$$

Oxidation half-reduction. Add 3 H_2O on the left side to balance the three O atoms on the right.

$$Br^- + 3H_2O \rightarrow BrO_3^-$$

ii) Next, balance the H atoms by adding H_2O on the side needing the H atoms and then adding the same number of OH^- ions to the other side of the equation. The net result is an addition of an H atom to the side to which we have added H_2O. When we add

$$\ldots OH^- \ldots \rightarrow \ldots H_2O \ldots$$

we are effectively adding one H atom to the right; when we add

$$\ldots H_2O \ldots \rightarrow \ldots OH^- \ldots$$

we are effectively adding an H atom to the left. Just remember that the number of OH^- that needs to be added is the same as the number of H_2O.

Reduction half-reaction. Add $4H_2O$ on the left and $4OH^-$ on the right to balance the four H atoms on the right:

$$MnO_4^- + 4H_2O \rightarrow MnO_2 + 2H_2O + 4OH^-$$

Oxidation half-reduction. The equation needs $6H_2O$ on the right and $6OH^-$ on the left to balance the six H atoms on the left:

$$Br^- + 3H_2O + 6OH^- \rightarrow BrO_3^- + 6H_2O$$

iii) The final part of this step, we simplify the two half-reactions by canceling like species on each side:

Reduction half-reaction. $MnO_4^- + 2H_2O \rightarrow MnO_2 + 4\ OH^-$

Oxidation half-reduction. $Br^- + 6OH^- \rightarrow BrO_3^- + 3H_2O$

4. Balance the electric charge by adding electrons.

Reduction half-reaction. To balance the 4– charge on the right and the 1– charge on the left, we need to add three electrons to the left:

$$MnO_4^- + 2H_2O + 3e^- \rightarrow MnO_2 + 4OH^-$$

Oxidation half-reduction. To balance 7– charge on the left and the 1– charge on the right, we need to add six electrons to the right:

$$Br^- + 6OH^- \rightarrow BrO_3^- + 3H_2O + 6e^-$$

5. Prepare the two half reactions for summation by making the number of electrons the same in both.

Multiply the reduction have equation by 2:

Reduction half-reactions: $2MnO_4^- + 4H_2O + 6e^- \rightarrow 2MnO_2 + 8OH^-$

Oxidation half-reduction: $Br^- + 6OH^- \rightarrow BrO_3^- + 3H_2O + 6e^-$

6. Combine the two half-reactions.

i) First, add the equations of the two half-reactions as they are written:

$$2MnO_4^- + 4H_2O + 6e^- + Br^- + 6OH^- \rightarrow 2MnO_2 + 8OH^- + BrO_3^- + 3H_2O + 6e^-$$

ii) Next, simplify the balanced equation by canceling like species on both sides of the equation. In this example, $6e^-$, $3H_2O$, and $6OH^-$ can be canceled:

$$2MnO_4^- + H_2O + Br^- \rightarrow 2MnO_2 + 2OH^- + BrO_3^-$$

iii) Finally, insert the state of each species:

$$2MnO_{4\,(aq)}^- + H_2O_{(l)} + Br^-_{(aq)} \rightarrow 2MnO_{2\,(s)} + 2OH^-_{(aq)} + BrO^-_{3\,(aq)}$$

c) This and the rest show only the essential steps, the steps you would be expected to show if asked to balance redox equation by the half-equation method.

The skeletal equation is

$$\begin{array}{ccccccc} 0 & \xrightarrow{\text{Oxidation}} & & & 2+ & & \\ \downarrow & & & & \downarrow & & \\ Cu_{(s)} & + & NO^-_{3(aq)} & + H^+_{(aq)} & \rightarrow Cu^{2+}_{(s)} & + & NO_{(g)} \\ & & \uparrow & & & & \uparrow \\ & & 5+ & \xrightarrow{\text{Reduction}} & & & 2+ \end{array}$$

Oxidizing agent: nitric acid (nitrogen undergoes reduction in oxidation number). Reducing agent copper (which undergoes increase in oxidation number).

Reduction half-equation: $NO_3^- \rightarrow NO$

Balanced reduction half-equation: $NO_3^- + 4H^+ + 3e^- \rightarrow NO + 2H_2O$

Oxidaton half-equation: $Cu \rightarrow Cu^{2+}$

Balanced oxidation half-equation: $Cu \rightarrow Cu^{2+} + 2e^-$

Multiplying the balanced reduction half-equation by 2 and the oxidation half-equation by 3 (to equate the number of electrons lost and gained), summing them, and adding state symbols:

$$2NO_3^- + 8H^+ + 6e^- \quad \rightarrow \quad 2NO + 4H_2O$$

$$3Cu \rightarrow 3Cu^{2+} + 6e^-$$

$$3\,Cu_{(s)} + 2NO_3^-{}_{(aq)} + 8H^+{}_{(aq)} \rightarrow 3Cu^{2+}{}_{(aq)} + 2NO_{(g)} + 4H_2O_{(l)}$$

d) The skeletal equation is

$$\begin{array}{c} 0 \xrightarrow{\text{Reduction}} 1- \\ \downarrow \qquad\qquad \downarrow \\ Cl_{2\,(g)} + S_2O_3^{2-}{}_{(aq)} \rightarrow Cl^-{}_{(aq)} + SO_4^{2-}{}_{(aq)} \\ \uparrow \qquad\qquad \uparrow \\ 2+ \xrightarrow{\text{Oxidation}} 6+ \end{array}$$

Oxidizing agent: chlorine. Reducing agent: thiosulfate.

Reduction half-equation: $Cl_2 \rightarrow Cl^-$

Balanced reduction half-equation: $Cl_2 + 2e^- \rightarrow 2\,Cl^-$

Oxidaton half-equation: $S_2O_3^{2-} \rightarrow SO_4^{2-}$

Balanced oxidation half-equation: $S_2O_3^{2-} + 5H_2O \rightarrow 2SO_4^{2-} + 10H^+ + 6e^-$

Multiplying the balanced reduction half-equation by 4 and the oxidation half-equation by 3, summing them, and adding state symbols:

$$4Cl_2 + 4e^- \rightarrow 8Cl^-$$

$$S_2O_3^{2-} + 5H_2O \rightarrow 2SO_4^{2-} + 10H^+ + 8e^-$$

$$4Cl_{2\,(g)} + S_2O_3^-{}_{(aq)} + 5\,H_2O_{(l)} \rightarrow 8Cl^-{}_{(aq)} + 8SO_4^-{}_{(g)} + 10H^+{}_{(l)}$$

e) The skeletal equation is

$$\begin{array}{c} 7+ \xrightarrow{\text{Reduction}} 2+ \\ \downarrow \qquad\qquad \downarrow \\ MnO_4^-{}_{(aq)} + H_2SO_{3(aq)} \rightarrow Mn^{2+}{}_{(s)} + HSO_4^-{}_{(aq)} \\ \uparrow \qquad\qquad \uparrow \\ 4+ \xrightarrow{\text{Oxidation}} 6+ \end{array}$$

Oxidizing agent: chlorine. Reducing agent: thiosulfate.

Reduction half-reaction: $MnO_4^- \rightarrow Mn^{2+}$

Balanced reduction half-equation: $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$

Oxidation half-equation: $H_2SO_3 \rightarrow HSO_4^-$

Balanced oxidation half-equation: $H_2SO_3 + H_2O \rightarrow HSO_4^- + 3H^+ + 2e^-$

Multiplying the balanced reduction half-equation by 2 and the oxidation half-equation by 5, summing them, and adding state symbols:

$$2MnO_4^- + 16H^+ + 10e^- \rightarrow 2Mn^{2+} + 8H_2O$$

$$5H_2SO_3 + 5H_2O \rightarrow 5HSO_4^- + 15H^+ + 10e^-$$

$$2MnO_{4\,(aq)}^- + H^+_{(aq)} + 5H_2SO_{3\,(aq)} \rightarrow 2Mn^{2+}_{(aq)} + 5HSO_{4\,(aq)}^- + 3H_2O_{(l)}$$

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f) The skeletal equation is

$$
\begin{array}{l}
0 \xrightarrow{\text{Reduction}} 1- \\
\downarrow \qquad\qquad\quad \downarrow \\
Cl_{2\,(g)} + H_2S_{(aq)} \rightarrow Cl^-_{(aq)} + S_{(s)} \\
\qquad\quad \uparrow \qquad\qquad\quad \uparrow \\
\qquad\quad 2- \xrightarrow{\text{Oxidation}} 0
\end{array}
$$

Oxidizing agent is chlorine, and reducing agent is hydrosulfuric acid.

The balanced equation is $Cl_{2\,(g)} + H_2S_{(aq)} \rightarrow 2Cl^-_{(aq)} + S_{(s)} + 2H^+_{(aq)}$

g) The skeletal equation is

$$
\begin{array}{l}
0 \xrightarrow{\text{Reduction}} 1- \\
\downarrow \qquad\qquad \downarrow \\
Br_{2\,(l)} \rightarrow Br^-_{(aq)} + BrO^-_{3(aq)} \\
\uparrow \qquad\qquad\qquad\qquad \uparrow \\
0 \xrightarrow{\text{Oxidation}} \quad 5+
\end{array}
$$

In this reaction, both the oxidizing and reducing agent is Br_2. This kind of reaction where the same substance undergoes both reduction and oxidation is called disproportionation.

Reduction half-equation: $Br_2 \rightarrow Br^-$

Balanced reduction half-equation: $Br_2 + 2e^- \rightarrow 2Br^-$

Oxidation half-equation: $Br_2 \rightarrow BrO_3^-$

Balanced oxidation half-equation: $Br_2 + 6H_2O + 12OH^- \rightarrow 2BrO_3^- + 12H_2O + 10e^-$

Multiplying the balanced reduction half-equation by 5, summing them, and adding state symbols:

$$5Br_2 + 10e^- \rightarrow 10Br^-$$

$$Br_2 + 6H_2O + 12OH^- \rightarrow 2BrO_3^- + 12H_2O + 10e^-$$

$$6Br_{2\,(l)} + 12OH^-_{(aq)} \rightarrow 10Br^-_{(aq)} + 2BrO^-_{3\,(aq)} + 6H_2O_{(l)}$$

---

h) The skeletal equation is

$$
\begin{array}{l}
0 \xrightarrow{\text{Oxidation}} \quad 6+ \\
\downarrow \qquad\qquad\qquad \downarrow \\
Cr^{3+}_{(aq)} + MnO_{2(s)} \rightarrow CrO^{2-}_{4\,(aq)} + Mn^{2+}_{(aq)} \\
\qquad\quad \uparrow \qquad\qquad\qquad\quad \uparrow \\
\qquad\quad 4+ \xrightarrow{\text{Reduction}} \quad 2+
\end{array}
$$

Oxidizing agent is manganese(IV) oxide, and reducing agent is chromium(III) ion.

The balanced equation is: $2Cr^{3+}{}_{(aq)} + 4OH^{-}{}_{(aq)} + 3MnO_{2\,(s)} \rightarrow 2CrO_4^{2-}{}_{(aq)} + 2H_2O_{(l)} + 3Mn^{2+}{}_{(aq)}$