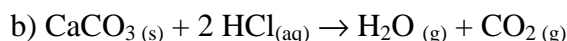


## First-Year Chemistry

### 2. Stoichiometry and Reactions

Solutions to further problems:

1. a)  $\text{CaCO}_3$  = solid,  $\text{HCl}$  = aqueous,  $\text{CaCl}_2$  = aqueous,  $\text{CO}_2$  = gas, and  $\text{H}_2\text{O}$  = liquid



c)  $0.500 \text{ mol CaCO}_3 \left( \frac{1 \text{ mol CO}_2}{1 \text{ mol CaCO}_3} \right) = 0.500 \text{ mol of CO}_2$  [3 sig figs]

d)  $0.250 \text{ mol CaCO}_3 \left( \frac{1 \text{ mol H}_2\text{O}}{1 \text{ mol CaCO}_3} \right) \left( \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} \right) = 4.50 \text{ g of H}_2\text{O}$  [3 sig figs]

e)  $10.5 \text{ g CaCO}_3 \left( \frac{1 \text{ mol CaCO}_3}{100.09 \text{ g CaCO}_3} \right) \left( \frac{1 \text{ mol CO}_2}{1 \text{ mol CaCO}_3} \right) \left( \frac{44.02 \text{ g CO}_2}{1 \text{ mol CO}_2} \right) = 4.62 \text{ g of CO}_2$  [3 sig figs]

f)  $20.0 \text{ g CaCO}_3 \left( \frac{1 \text{ mol CaCO}_3}{100.09 \text{ g CaCO}_3} \right) \left( \frac{1 \text{ mol H}_2\text{O}}{1 \text{ mol CaCO}_3} \right) \left( \frac{6.02 \times 10^{23} \text{ molecules H}_2\text{O}}{1 \text{ mol H}_2\text{O}} \right) = 1.20 \times 10^{23}$   
molecules of water [3 sig  
figs]

g)  $14.0 \text{ g CO}_2 \left( \frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} \right) \left( \frac{2 \text{ mol HCl}}{1 \text{ mol CO}_2} \right) = 0.636 \text{ mol HCl}$  [3 sig figs]

2.  $(13.7 \text{ cm}^3 \text{ H}_2\text{SO}_4) \left( \frac{1.20 \text{ mol H}_2\text{SO}_4}{1000 \text{ cm}^3 \text{ H}_2\text{SO}_4} \right) \left( \frac{1}{0.100 \text{ dm}^3} \right) = 0.1644 = 0.164 \text{ M}$  (3 sig figs)

3. (a)  $[\text{NH}_4^+] = \left( \frac{1.20 \text{ mol (NH}_4)_2\text{SO}_4}{1 \text{ dm}^3} \right) \left( \frac{2 \text{ mol NH}_4^+}{1 \text{ mol (NH}_4)_2\text{SO}_4} \right) = 2.40 \text{ M}$  (3 sig figs)

$$[\text{SO}_4^{2-}] = \left( \frac{1.20 \text{ mol (NH}_4)_2\text{SO}_4}{1 \text{ dm}^3} \right) \left( \frac{1 \text{ mol SO}_4^{2-}}{1 \text{ mol (NH}_4)_2\text{SO}_4} \right) = 1.20 \text{ M}$$
 (3 sig figs)

(b)  $[\text{Ca}^{2+}] = \left( \frac{0.250 \text{ mol CaCl}_2}{1 \text{ dm}^3} \right) \left( \frac{1 \text{ mol Ca}^{2+}}{1 \text{ mol CaCl}_2} \right) = 0.250 \text{ M}$  (3 sig figs)

$$[\text{Cl}^-] = \left( \frac{0.250 \text{ mol CaCl}_2}{1 \text{ dm}^3} \right) \left( \frac{2 \text{ mol Cl}^-}{1 \text{ mol CaCl}_2} \right) = 0.500 \text{ M}$$
 (3 sig figs)

(c)  $[\text{Fe}^{2+}] = \left( \frac{0.400 \text{ mol Fe}_2(\text{SO}_4)_3}{1 \text{ dm}^3} \right) \left( \frac{2 \text{ mol Fe}^{2+}}{1 \text{ mol Fe}_2(\text{SO}_4)_3} \right) = 0.800 \text{ M}$  (3 sig figs)

$$[\text{SO}_4^{2-}] = \left( \frac{0.400 \text{ mol Fe}_2(\text{SO}_4)_3}{1 \text{ dm}^3} \right) \left( \frac{3 \text{ mol SO}_4^{2-}}{1 \text{ mol Fe}_2(\text{SO}_4)_3} \right) = 1.20 \text{ M}$$
 (3 sig figs)

4. The equation for the reaction is  $2\text{HClO}_3(\text{aq}) + \text{Ba}(\text{OH})_2(\text{aq}) \rightarrow \text{Ba}(\text{ClO}_3)_2(\text{aq}) + 2\text{H}_2\text{O}(\text{l})$   
Answer is 0.100 (same as the number of moles of  $\text{HClO}_3$  since firstly it's the limiting reagent, and secondly, the ratio of  $\text{HClO}_3$  to  $\text{H}_2\text{O}$  are equal).

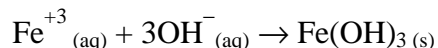
5. Equation is:  $\text{AgNO}_3(\text{aq}) + \text{KBr}(\text{aq}) \rightarrow \text{AgBr}(\text{s}) + \text{KNO}_3(\text{aq})$

$$(2.00 \text{ g AgNO}_3) \left( \frac{1 \text{ mol AgNO}_3}{169.88 \text{ g AgNO}_3} \right) \left( \frac{1 \text{ mol AgBr}}{1 \text{ mol AgNO}_3} \right) \left( \frac{187.77 \text{ g AgBr}}{1 \text{ mol AgBr}} \right) = 2.210 \text{ g} = 2.21 \text{ g AgBr}$$

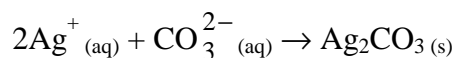
$$(4.00 \text{ g KBr}) \left( \frac{1 \text{ mol KBr}}{119.00 \text{ g KBr}} \right) \left( \frac{1 \text{ mol AgBr}}{1 \text{ mol KBr}} \right) \left( \frac{187.77 \text{ g AgBr}}{1 \text{ mol AgBr}} \right) = 6.311 \text{ g} = 6.31 \text{ g AgBr}$$

AgNO<sub>3</sub> 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)



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

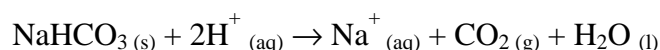


- c) No, a reaction does not take place; all acetates and nitrates are soluble.

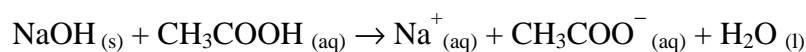
- d) Yes.



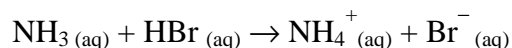
- e) Yes.



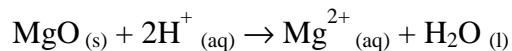
- f) Yes.



- g) Yes.



- h) Yes.



7. a) Total charge on ion = -1

$$\Rightarrow \text{charge on I} + 3(\text{charge on O}) = -1$$

$$\Rightarrow \text{charge on I} + 3(-2) = -1$$

$$\Rightarrow \text{charge on I} - 6 = -1$$

$$\Rightarrow \text{charge on I} = -1 + 6$$

$$\Rightarrow \text{charge on I} = 5$$

Oxidation number of I in iodate ion = 5+.

- b) Total charge on compound = 0

$$\Rightarrow \text{charge on H} + \text{charge on Cl} + \text{charge on O} = 0$$

$$\Rightarrow 1 + \text{charge on Cl} + (-2) = 0$$

$$\Rightarrow 1 + \text{charge on Cl} - 2 = 0$$

$$\Rightarrow \text{charge on Cl} = 2 - 1$$

$$\Rightarrow \text{charge on Cl} = 1$$

Oxidation number of Cl in HClO = 1+

- c) Oxidation number of N in NO = 2+

- d) Oxidation number of N in HNO<sub>3</sub> = 5+

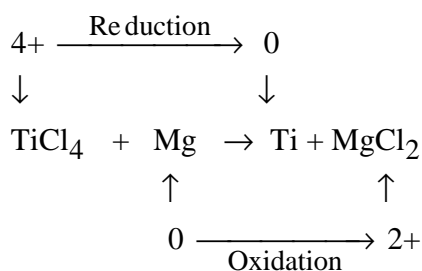
- e) Oxidation number of Mn in MnO<sub>4</sub><sup>-</sup> = 7+

- f) Oxidation number of S in S<sub>2</sub>O<sub>3</sub><sup>2-</sup> = 2+

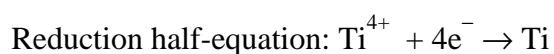
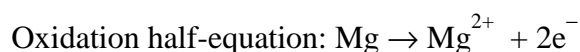
g) Oxidation number of S in  $\text{SO}_4^{2-} = 6+$

h) Oxidation number of Mn in  $\text{MnO}_4^- = 6+$

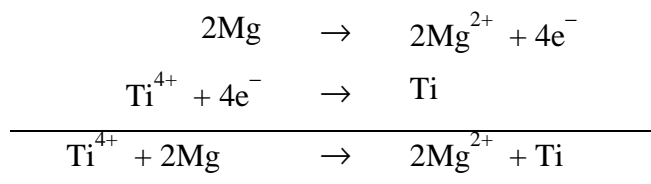
8. a)



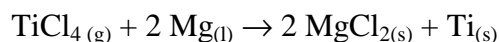
Balancing each half-equation:



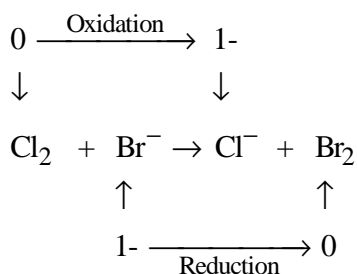
Multiplying the oxidation half equation by 2 and adding the two balanced half-equations:



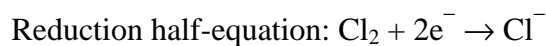
Adding and balancing the rest of the elements:



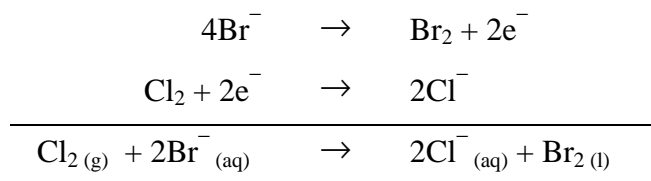
b)



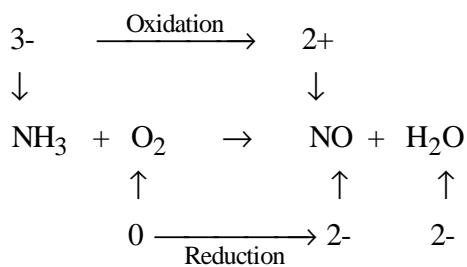
Balancing each half-equation:



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



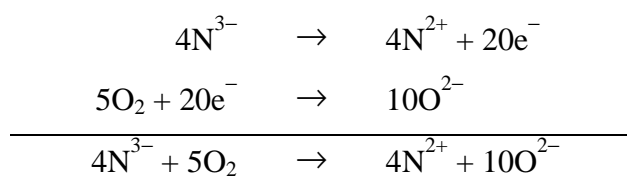
c)



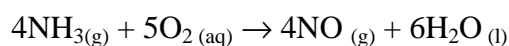
Balancing each half-equation:



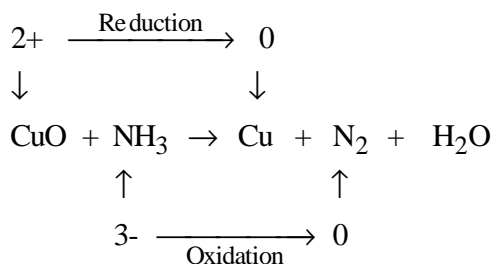
Multiplying the oxidation half-equation by 4 and the reduction half-equation by 5, and adding the two balanced half-equations:



Adding and balancing the rest of the elements:



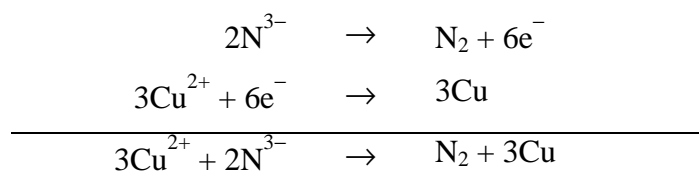
d)



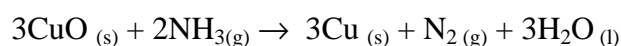
Balancing each half-equation:



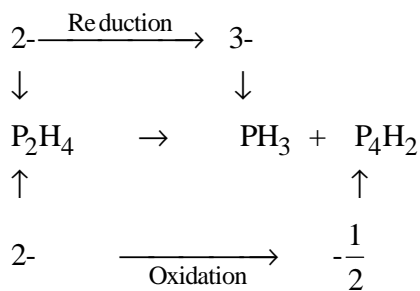
Multiplying the reduction half-equation by 3 and adding the two balanced half-equations:



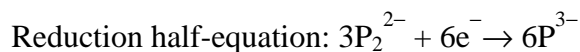
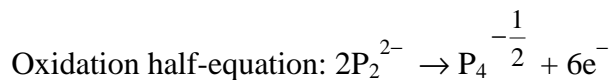
Adding and balancing the rest of the elements



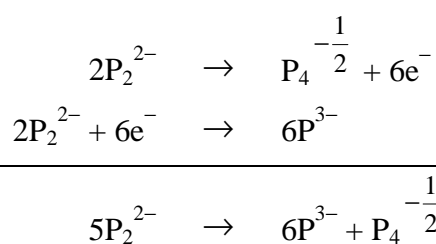
e)



Balancing each half-equation:



Adding the two balanced half-equations:

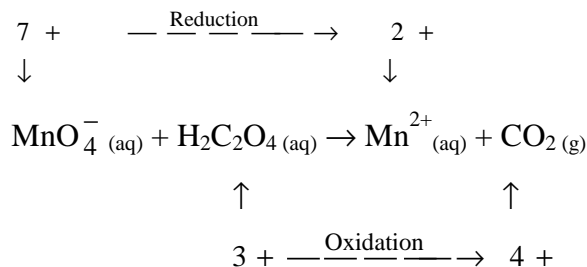


Adding and balancing the rest of the elements:



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

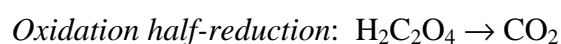
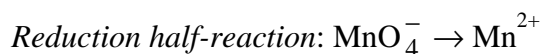
a) Half-equation method: First, the skeletal equation:



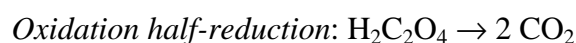
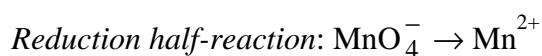
Oxidation number of manganese decreases from 7+ to 2+, hence  $\text{MnO}_4^{2-}$  is the oxidizing agent.

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.



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



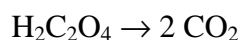
3. Balance the oxygen and hydrogen atoms. For oxanions in *acidic* solution, we balance H and O by adding  $\text{H}^+$  and  $\text{H}_2\text{O}$ :

i) First balance the O atoms by adding H<sub>2</sub>O:

*Reduction half-reaction.* Add 4 H<sub>2</sub>O on the right to balance the four O atoms on the left:

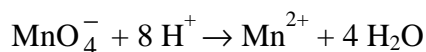


*Oxidation half reaction:* The O atoms are already balanced.

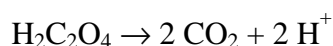


ii) Next, add H<sup>+</sup> to balance the H atoms on either side of the half reactions:

*Reduction half-reaction:* Add 8 H<sup>+</sup> on the left to balance the eight H atoms on the right:

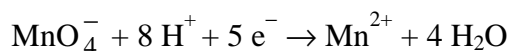


*Oxidation half reaction.* The equation needs 2 H<sup>+</sup> on the right to balance the H atoms in the H<sub>2</sub>C<sub>2</sub>O<sub>2</sub> molecule on the left.



4. Balance the electric charges by adding electrons.

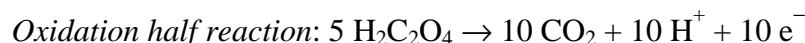
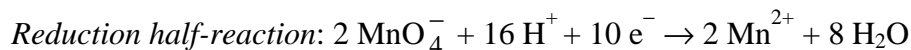
*Reduction half-reaction.* To balance the 2+ charge on the right and 7+ charge on the left, we need to add five electrons to the left:



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

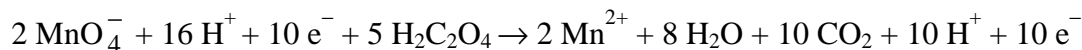


5. prepare the two half reactions for summation by making the number of electrons the same in both. (Electrons cannot be created or destroyed.) Multiply the reduction half-reaction by 2 and the oxidation half-reaction by 5:

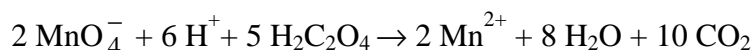


6. Combine the two half-reactions.

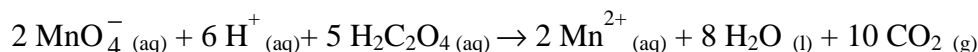
i) First, add the two half-reactions in step 5:



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

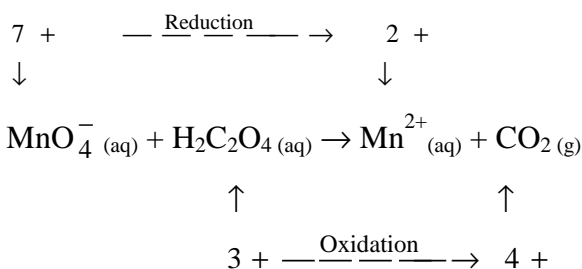


iii) Finally, insert the state of each species:



The expression is the fully balanced net ionic equation.

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



Step 1: Balancing the total change in oxidation numbers

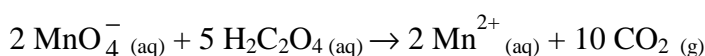
Total change in oxidation number of manganese = final – initial = 2 – 7 = – 5

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



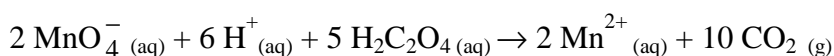
Step 2: Balancing charges by adding  $\text{H}^+$ .

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  $6\text{H}^+$ .



Step 3: Balancing excess hydrogen and oxygen by adding  $\text{H}_2\text{O}$ .

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

Number of H atoms on the right hand side = 0

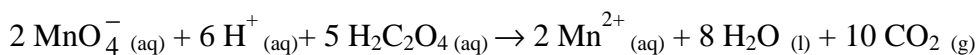
Right hand requires 16 H atoms.

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

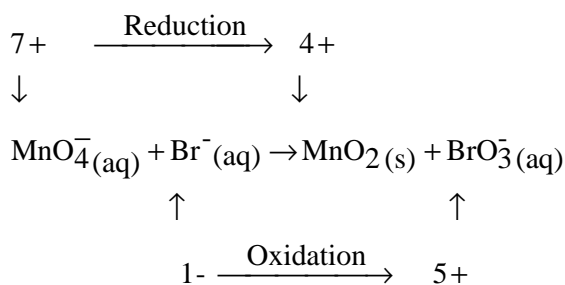
Number of O atoms on the right hand side = 10(2) = 20

Right hand requires 8 O atoms.

Therefore, right hand side requires  $8\text{H}_2\text{O}$  (16 H and 8 O atoms)



b) The skeletal equation:

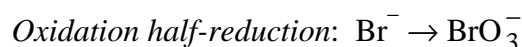
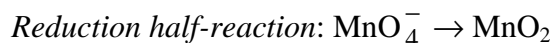


Since the oxidation number of manganese decreases from 7+ to 4+ it is the oxidizing agent.

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.



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.

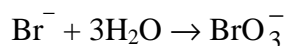
3. Balance the oxygen and hydrogen atoms. For oxoanions in *basic* solution, we balance H and O by adding  $\text{OH}^-$  and  $\text{H}_2\text{O}$ .

i) First balance the O atoms by adding  $\text{H}_2\text{O}$ :

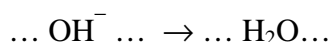
*Reduction half-reaction.* Add 2  $\text{H}_2\text{O}$  on the right side to balance the two extra O atoms on the left.



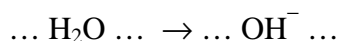
*Oxidation half-reduction.* Add 3  $\text{H}_2\text{O}$  on the left side to balance the three O atoms on the right.



ii) Next, balance the H atoms by adding  $\text{H}_2\text{O}$  on the side needing the H atoms and then adding the same number of  $\text{OH}^-$  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  $\text{H}_2\text{O}$ . When we add



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



we are effectively adding an H atom to the left. Just remember that the number of  $\text{OH}^-$  that needs to be added is the same as the number of  $\text{H}_2\text{O}$ .

*Reduction half-reaction.* Add 4 $\text{H}_2\text{O}$  on the left and 4 $\text{OH}^-$  on the right to balance the four H atoms on the right:



*Oxidation half-reduction.* The equation needs 6 $\text{H}_2\text{O}$  on the right and 6 $\text{OH}^-$  on the left to balance the six H atoms on the left:



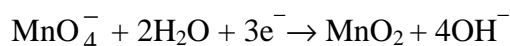
iii) The final part of this step, we simplify the two half-reactions by canceling like species on each side:



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:



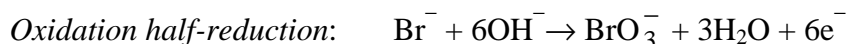
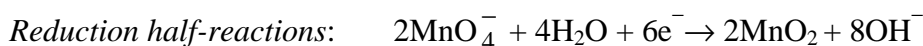


*Oxidation half-reduction.* To balance 7- charge on the left and the 1- charge on the right, we need to add six electrons to the right:



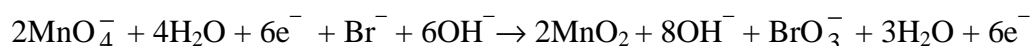
5. Prepare the two half reactions for summation by making the number of electrons the same in both.

Multiply the reduction half equation by 2:

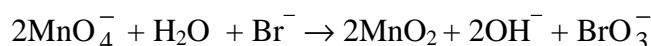


6. Combine the two half-reactions.

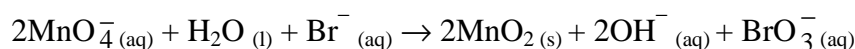
i) First, add the equations of the two half-reactions as they are written:



ii) Next, simplify the balanced equation by canceling like species on both sides of the equation. In this example,  $6\text{e}^-$ ,  $3\text{H}_2\text{O}$ , and  $6\text{OH}^-$  can be canceled:

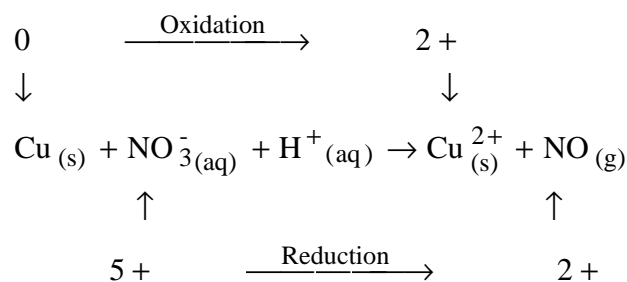


iii) Finally, insert the state of each species:

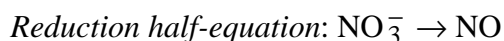


c) This and the rest show only the essential steps, the steps you would be expected to show if asked to balance redox equation by the half-equation method.

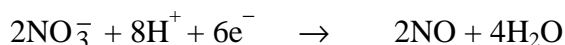
The skeletal equation is

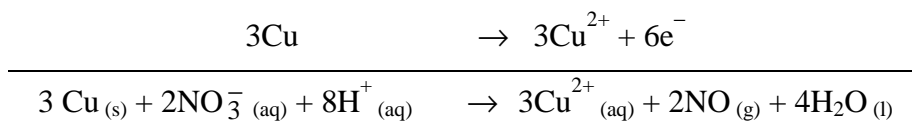


Oxidizing agent: nitric acid (nitrogen undergoes reduction in oxidation number). Reducing agent copper (which undergoes increase in oxidation number).

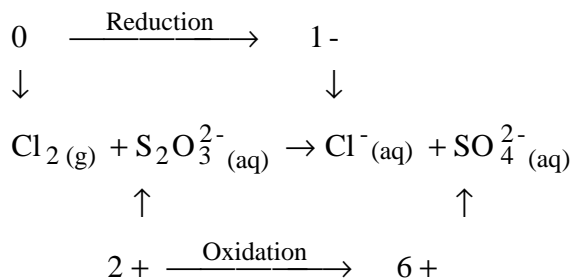


Multiplying the balanced reduction half-equation by 2 and the oxidation half-equation by 3 (to equate the number of electrons lost and gained), summing them, and adding state symbols:





d) The skeletal equation is



Oxidizing agent: chlorine. Reducing agent: thiosulfate.

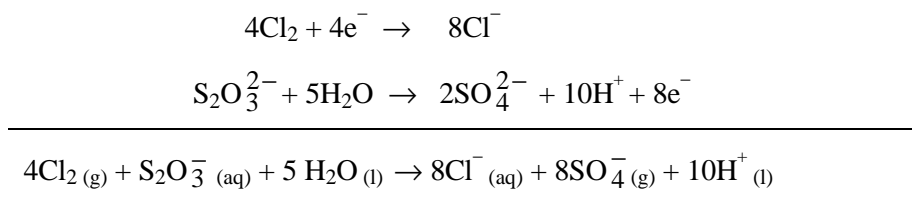
*Reduction half-equation:*  $\text{Cl}_2 \rightarrow \text{Cl}^{-}$

*Balanced reduction half-equation:*  $\text{Cl}_2 + 2\text{e}^{-} \rightarrow 2\text{Cl}^{-}$

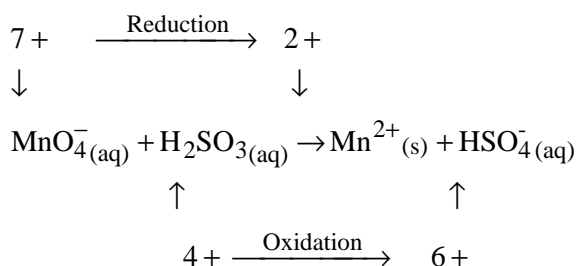
*Oxidation half-equation:*  $\text{S}_2\text{O}_3^{2-} \rightarrow \text{SO}_4^{2-}$

*Balanced oxidation half-equation:*  $\text{S}_2\text{O}_3^{2-} + 5\text{H}_2\text{O} \rightarrow 2\text{SO}_4^{2-} + 10\text{H}^{+} + 6\text{e}^{-}$

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



e) The skeletal equation is



Oxidizing agent: chlorine. Reducing agent: thiosulfate.

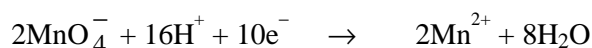
*Reduction half-reaction:*  $\text{MnO}_4^{-} \rightarrow \text{Mn}^{2+}$

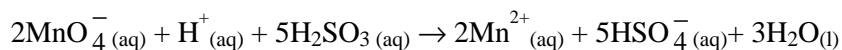
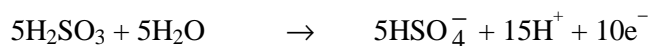
*Balanced reduction half-equation:*  $\text{MnO}_4^{-} + 8\text{H}^{+} + 5\text{e}^{-} \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$

*Oxidation half-equation:*  $\text{H}_2\text{SO}_3 \rightarrow \text{HSO}_4^{-}$

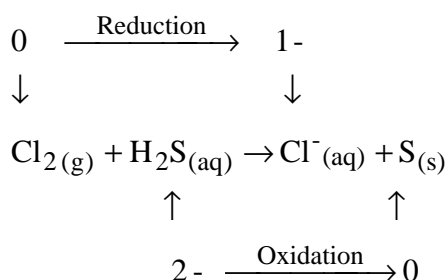
*Balanced oxidation half-equation:*  $\text{H}_2\text{SO}_3 + \text{H}_2\text{O} \rightarrow \text{HSO}_4^{-} + 3\text{H}^{+} + 2\text{e}^{-}$

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

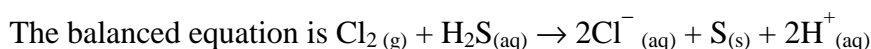




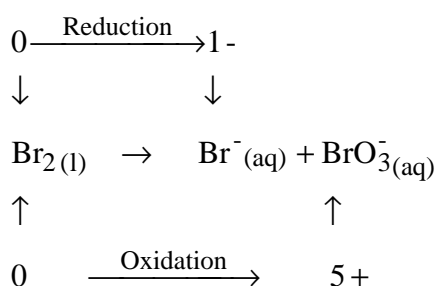
f) The skeletal equation is



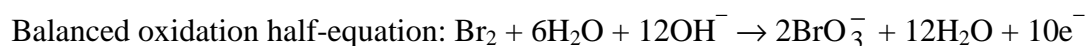
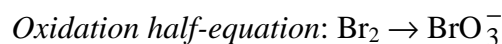
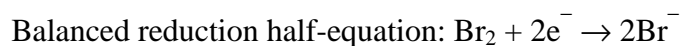
Oxidizing agent is chlorine, and reducing agent is hydrosulfuric acid.



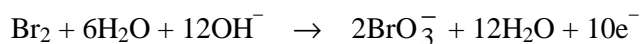
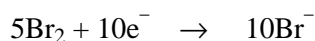
g) The skeletal equation is



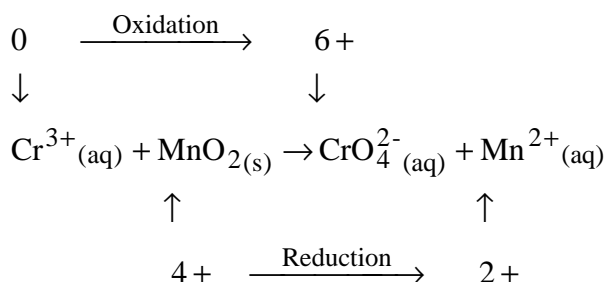
In this reaction, both the oxidizing and reducing agent is  $\text{Br}_2$ . This kind of reaction where the same substance undergoes both reduction and oxidation is called disproportionation.



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



h) The skeletal equation is



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

