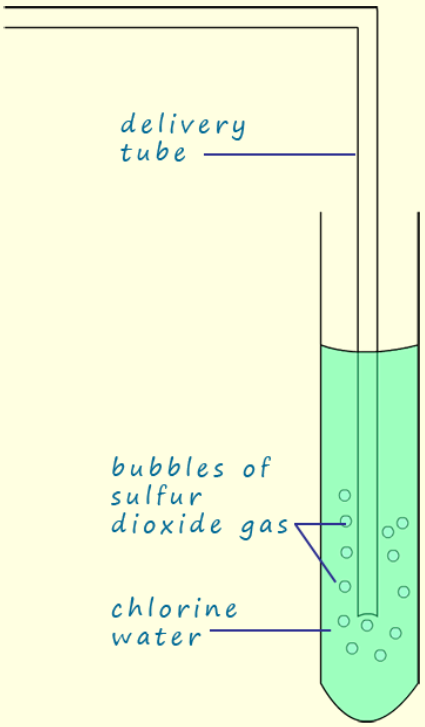
1. Sulfur dioxide gas ($SO_{2(g)}$) was bubbled through a solution of chlorine water ($Cl_{2(aq)}$). The chlorine was reduced to form chloride ions and the sulfur dioxide is oxidised to form a solution of sulfuric acid (H_2SO_4).



a. Write a half-equation to show the reduction of chlorine to form aqueous chloride ions.

ii. Explain why this reaction can be described as reduction.

b. In this reaction the sulfur dioxide gas is oxidised to form sulfate ions.

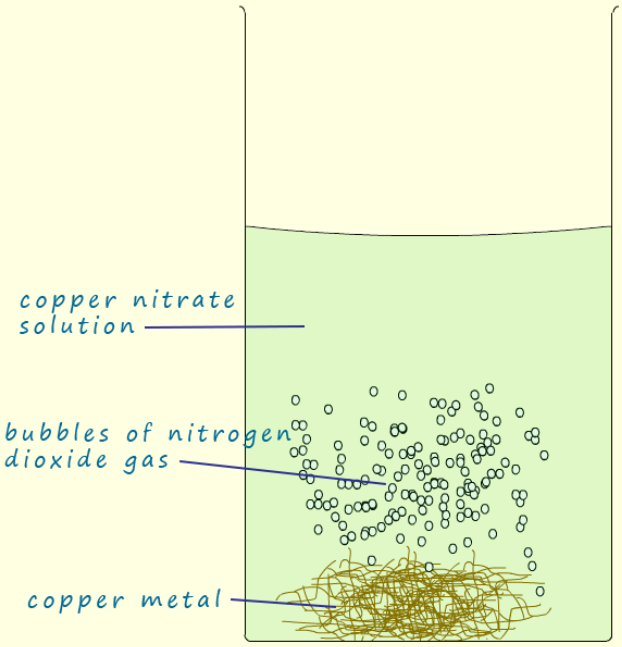
i. What is the oxidation state of the sulfur atom in $SO_{2(g)}$ and in the sulfate ion (SO_4^{2-})?

ii. Explain why this reaction is an oxidation reaction.

c. Write a half-equation to show the oxidation of $SO_{2(g)}$ to form the sulfate ion (SO_4^{2-}).

d. Combine the 2 half-equations to give the complete equation for this redox reaction.

2. Nitric acid (HNO_3) is an oxidising acid. It will react with unreactive metals such as copper to produce colourless nitrogen monoxide gas (NO) and a solution of copper (II) nitrate.



a. Write a half-equation for the oxidation of copper to form copper (II) ions.

b. What is the oxidation state of nitrogen in nitrate and nitrogen monoxide?

i. Write a half-equation to show the reduction of nitrate ions (NO_3^-) to form the colourless gas nitrogen monoxide.

ii. Combine the 2 half-equations to produce an equation for the overall redox reaction (take care to ensure that the electrons balance in this reaction).

3. In acidic conditions hydrogen peroxide (H_2O_2) is reduced to water by iodide ions (I^-), the iodide is oxidised to iodine.

i. What is the oxidation state of oxygen in hydrogen peroxide?

ii. What is the oxidation state of iodine in an iodide ion?

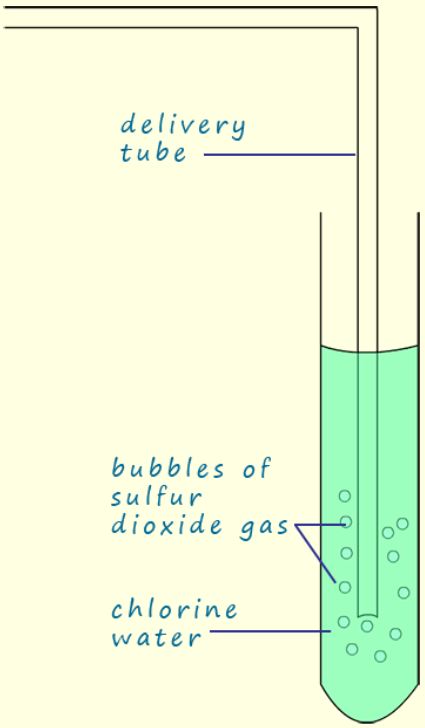
b. Write a half-equation to show the oxidation of iodide ions to form iodine.

c. Write a half-equation to show the rduction of hydrogen perxide in acidic conditions to form water.

d. Construct an equation from the two half-equations to show the overall reaction taking place.

Answers

1. Sulfur dioxide gas ($SO_{2(g)}$) was bubbled through a solution of chlorine water ($Cl_{2(aq)}$). The chlorine was reduced to form chloride ions and the sulfur dioxide is oxidised to form a solution of sulfuric acid (H_2SO_4).



a. Write a half-equation to show the reduction of chlorine to form aqueous chloride ions.

$Cl_{2(aq)} + 2e \longrightarrow 2Cl^-_{(aq)}$

ii. Explain why this reaction can be described as reduction.

The chlorine gains 2 electrons

b. In this reaction the sulfur dioxide gas is oxidised to form sulfate ions.

i. What is the oxidation state of the sulfur atom in $SO_{2(g)}$ and in the sulfate ion (SO_4^{2-})?

$SO_{2(g)}$ S=+4 O=-2X2=-4 SO_4^{2-} S=+6 O=-2x4=-8 so +6-8=-2 charge on the ion.

ii. Explain why this reaction is an oxidation reaction.

The oxidation state of the sulfur atom increases from + 4 to +6, that is it is losing 2 electrons.

c. Write a half-equation to show the oxidation of $SO_{2(g)}$ to form the sulfate ion (SO_4^{2-}).

$SO_{2(g)} + 2H_2O_{(l)} \longrightarrow SO_4^{2-}{}_{(aq)} + 4H^+_{(aq)} + 2e$

d. Combine the 2 half-equations to give the complete equation for this redox reaction.

$Cl_{2(aq)} + 2e \longrightarrow 2Cl^-_{(aq)}$

$SO_{2(g)} + 2H_2O \longrightarrow SO_4^{2-}{}_{(aq)} + 4H^+_{(aq)} + \cancel{2e}$

$SO_{2(g)} + 2H_2O_{(l)} + Cl_{2(aq)} \longrightarrow SO_4^{2-}{}_{(aq)} + 4H^+_{(aq)} + 2Cl^-_{(aq)}$

Or

$SO_{2(g)} + 2H_2O_{(l)} + Cl_{2(aq)} \longrightarrow H_2SO_4^{2-}{}_{(aq)} + 2HCl^-_{(aq)}$

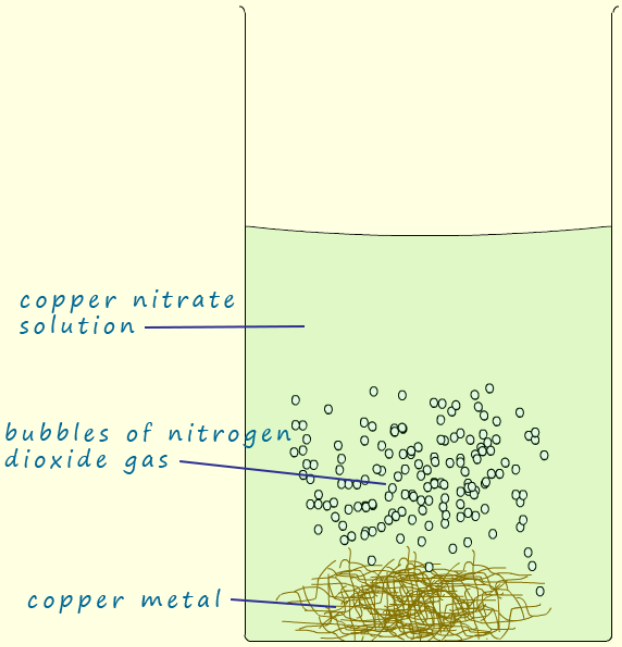
2. Nitric acid (HNO_3) is an oxidising acid. It will react with unreactive metals such as copper to produce colourless nitrogen monoxide gas (NO) and a solution of copper (II) nitrate.

a. Write a half-equation for the oxidation of copper to form copper (II) ions.

Oxidation : $Cu \longrightarrow Cu^{2+} + 2e$

b. What is the oxidation state of nitrogen in nitrate and nitrogen monoxide?

NO_3^- oxidation state of N is +5 NO oxidation state of nitrogen is +2



i. Write a half-equation to show the reduction of nitrate ions (NO_3^-) to form the colourless gas nitrogen monoxide.

Reduction : $NO_3^- \longrightarrow NO$

Add $2H_2O$ to the products side of the equation to balance off the oxygen atoms, this gives:

Reduction : $NO_3^- \longrightarrow NO + 2H_2O$

Now add $4H^+$ to the reactants side to balance off the hydroegn atoms, this gives:

Reduction : $NO_3^- + 4H^+ \longrightarrow NO + 2H_2O$

Total charge on reactant side of equation is 4+ -1 = +3

Total charge on products isde of equation is 0.

So to balance charges add 3e to reactants side of equation. This will give a total charge of 0 on both sides of the equation:

Reduction : $NO_3^- + 4H^+ + 3e \longrightarrow NO + 2H_2O$

ii. Combine the 2 half-equations to produce an equation for the overall redox reaction (take care to ensure that the electrons balance in this reaction).

Oxidation : $Cu \longrightarrow Cu^{2+} + 2e$

Reduction : $NO_3^- + 4H^+ + 3e \longrightarrow NO + 2H_2O$

The electrons do not balance, oxidation of copper releases 2e but reduction reaction requires 3e. So x3 the oxidation reaction and x2 the reduction reaction:

Oxidation : $3Cu \longrightarrow 3Cu^{2+} + 6e$

Reduction : $2NO_3^- + 8H^+ + 6e \longrightarrow 2NO + 4H_2O$

Combine two equations to get the overall redox equation:

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

The products of the reaction are Copper (II) nitrate and nitrogen monoxide gas. In the above equation the total charges on the reactants side and products side do not balance. The $8H^+$ ions are coming from the nitric acid (HNO_3), the product $3Cu^{2+}{}_{(aq)}$ is actually $3Cu(NO_3)_2$, this will require an additional 6 x NO_3^- ions. These nitrate ions are not being reduced so do not appear in the redox equation.

$$3Cu_{(s)} + 8NO_3^-{}_{(aq)} + 8H^+{}_{(aq)} \longrightarrow 3Cu^{2+}{}_{(aq)} + 6NO_3^-{}_{(aq)} + 2NO_{(g)} + 4H_2O_{(l)}$$

Or simply

$$3Cu_{(s)} + 8HNO_{3(aq)} \longrightarrow 3Cu(NO_3)_2{}_{(aq)} + 2NO_{(g)} + 4H_2O_{(l)}$$

3. In acidic conditions hydrogen peroxide (H_2O_2) is reduced to water by iodide ions (I^-), the iodide is oxidised to iodine.

i. What is the oxidation state of oxygen in hydrogen peroxide?

Oxidation of oxygen in H_2O_2 is -1

ii. What is the oxidation state of iodine in an iodide ion?

Oxidation state of iodine in iodide is -1

b. Write a half-equation to show the oxidation of iodide ions to form iodine.

$$2I^- \longrightarrow I_2 + 2e$$

c. Write a half-equation to show the rduction of hydrogen perxide in acidic conditions to form water.

$H_2O_2 \longrightarrow H_2O$

As before add oxygen atoms to the oxygen deficient side, in this case add water to the products side, this gives:

$H_2O_2 \longrightarrow 2H_2O$

Now balance off the H atoms, so add $2H^+$ to the reactants side

$H_2O_2 + 2H^+ \longrightarrow 2H_2O$

Now balance off the charges, need to add 2e to the reactants, this give a total charge of O on both sides of the equation.

$H_2O_2 + 2H^+ + 2e \longrightarrow 2H_2O$

d. Construct an equation from the two half-equations to show the overall reaction taking place.

$H_2O_2 + 2H^+ + 2e \longrightarrow 2H_2O$

$2I^- \longrightarrow I_2 + 2e$

The electrons for the oxidation and reduction reactions balance so simply add the two equations together to get the overall equation for the reaction.

$H_2O_2 + 2H^+ + 2I^- \longrightarrow 2H_2O + I_2$