A) Solid Iodine trichloride is made in two steps.
(1) Iodine is reacted with Chlorine gas to form Iodine monochloride
(2) Treatment with more chlorine

Write a balanced equation for each step, write the overall reaction, and then find how many grams of Iodine is needed to make 2.45kg of Iodine trichloride.

$$I_2 + Cl_2 \longrightarrow 2\,ICl$$

$$ICl + Cl_2 \longrightarrow ICl_3$$

$$I_2 + Cl_2 \longrightarrow 2\,ICl$$

$$2\,ICl + 2\,Cl_2 \longrightarrow 2\,ICl_3$$

Subtract the differences

$$I_2 + 3Cl_2 \longrightarrow 2\,ICl_3$$

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$$I_2 + 3Cl_2 \longrightarrow 2\ ICl_3$$

Plan, convert 2.45kg to mols, use stoichiometry to determine moles of Iodine required, convert to mass.

2.45kg = 2.45×10^3 g

Mols = mass / MW
Mols = 2.45×10^3 g / 233.2 g/mol
Mols = 10.506 mols ICl_3

Using stoichiometry, we see that every one mol of Iodine (I_2) produces 2 mols of (ICl_3). That means we need 10.506 / 2 mols = 5.253 mols of I_2

Now convert that to grams mass = mols x MW
mass = 5.253 x 253.8 g/mol
mass = 1.33×10^3 grams of I_2

Calculate the maximum amount (mols and mass) of iodic acid (HIO_3) that can be formed when 635g of Iodine trichloride reacts with 118.5g of water.

$$2\ ICl_3 + 3\ H_2O \longrightarrow ICl + HIO_3 + 5\ HCl$$

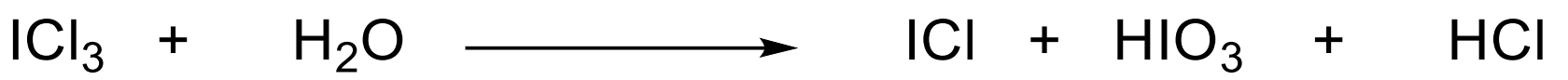
Plan, balance reaction. Find limiting reagent, determine theoretical yield using limiting reagent

Mols of ICl_3 = 635 g / 233.2 g/mol
Mols of ICl_3 = 2.722985 mols
Theoretical yield of HIO_3 = 2.722985 / 2
Theoretical yield of HIO_3 = 1.36 mols

Mols of water = 118.5 g / 18 g/mol
Mols of water = 6.57603 mols
Theoretical yield of HIO_3 = 6.57603 / 3
Theoretical yield of HIO_3 = 2.19 mols

ICl_3 is the limiting reagent!

Calculate the maximum amount (mols and mass) of iodic acid (HIO_3) that can be formed when 635g of Iodine trichloride reacts with 118.5g of water.

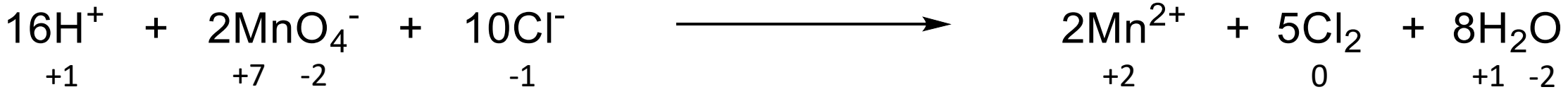


Plan, balance reaction. Find limiting reagent, determine theoretical yield using limiting reagent

Theoretical yield of HIO_3 = 1.36 mols
Theoretical yield of HIO_3 = 1.36 mols x 175.9 g/mol
Theoretical yield of HIO_3 = 239g

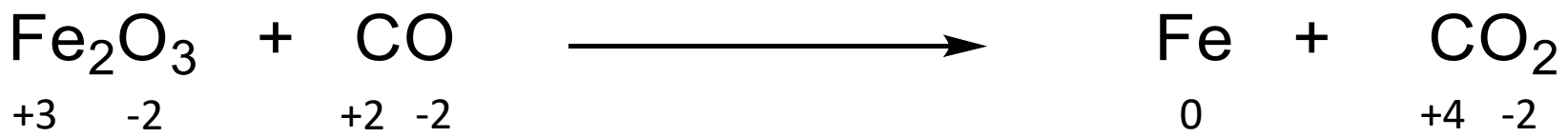
Consider the follow redox reaction.

A) Which species is being oxidized
B) Which species is being reduced
C) Which species is the oxidizing agent
D) Which species is the reducing agent
E) What species are the electrons traveling between and in which direction



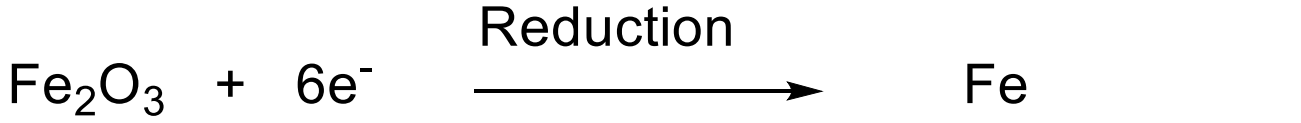
A) Oxidation means loss of electrons. Chlorine goes from -1 to 0, it got oxidized.
B) Reduction is the opposite of oxidation, it gains electrons. Manganese started as +7 and went to +2, it was reduced.
C) Oxidizing agents are chemicals that oxidize other chemicals. They STEAL electrons. The Manganese oxidized (stole electrons from) the chlorine
D) The reducing agent gives electrons to other chemicals, in this case chlorine reduced Manganese.
E) Chlorine GAVE electrons to Manganese.

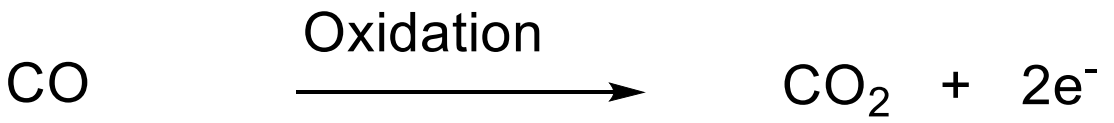
Balance the following redox reaction (under acidic conditions)



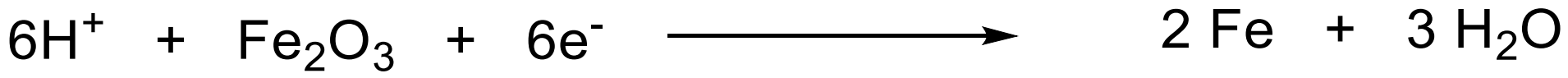
Plan, follow rules for balancing redox reactions under acidic conditions

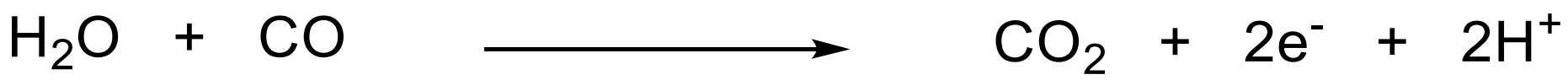
1) Assign charges, and write out half reactions.





Balance each reaction, starting with the reducing / oxidizing agents (everything but Oxygen and Hydrogen). Then balance Oxygen with water, and Hydrogen with H^+ (we're under acidic conditions after all)





Now balance the total equation by eliminating the electrons (e^-)

$$6H^+ + Fe_2O_3 + \boxed{6e^-} \longrightarrow 2\,Fe + 3\,H_2O$$

$$H_2O + CO \longrightarrow CO_2 + \boxed{2e^-} + 2H^+$$

Multiply the second reaction by 3

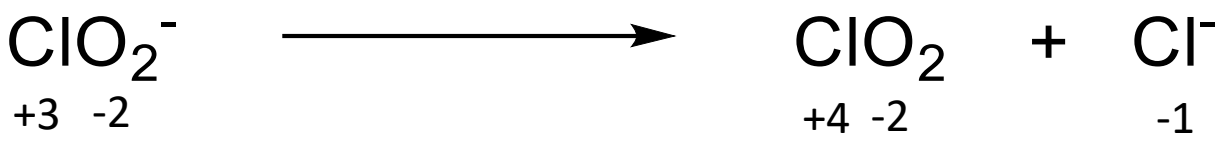
$$6H^+ + Fe_2O_3 + \boxed{6e^-} \longrightarrow 2\,Fe + 3\,H_2O$$

$$3\,H_2O + 3\,CO \longrightarrow 3\,CO_2 + \boxed{6\,e^-} + 6\,H^+$$

Subtract the differences

$$Fe_2O_3 + 3\,CO \longrightarrow 2\,Fe + 3\,CO_2$$

Balance the following redox reaction (under acidic conditions)



Plan, follow rules for balancing redox reactions under acidic conditions

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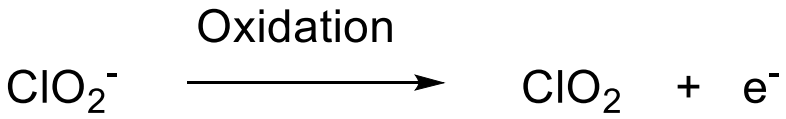
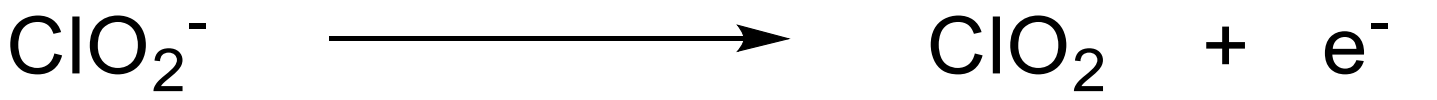
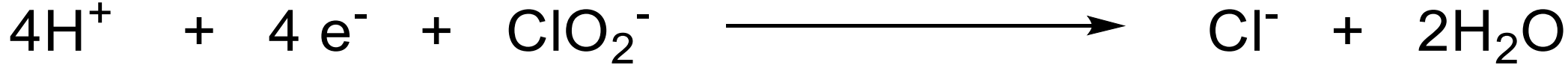


Balance each reaction, starting with the reducing / oxidizing agents. Then balance Oxygen with water, and Hydrogen with H^+ (we're under acidic conditions after all)





Now balance the total equation by eliminating the electrons (e^-)

$$4H^+ + 4\,e^- + ClO_2^- \longrightarrow Cl^- + 2H_2O$$

$$ClO_2^- \longrightarrow ClO_2 + e^-$$

Multiply the second reaction by 4

$$4H^+ + 4\,e^- + ClO_2^- \longrightarrow Cl^- + 2H_2O$$

$$4\,ClO_2^- \longrightarrow 4\,ClO_2 + 4\,e^-$$

Subtract the differences

$$4H^+ + 5ClO_2^- \longrightarrow 4\,ClO_2 + Cl^- + 2H_2O$$

Balance the following redox reaction (under basic conditions)

$$\underset{+3\ -2\ +1}{Cr(OH)_3} + \underset{0}{Br_2} \longrightarrow \underset{+6\ -2}{CrO_4^{2-}} + \underset{-1}{Br^{-1}}$$

Plan, follow rules for balancing redox reactions under BASIC conditions (the same except with one slight variation)

1) Assign charges, and write out half reactions.

$$Cr(OH)_3 \xrightarrow{\text{Oxidation}} CrO_4^{2-} + 3\ e^-$$

$$e^- + Br_2 \xrightarrow{\text{Reduction}} Br^-$$

Balance each reaction, starting with the reducing / oxidizing agents. Then balance Oxygen with water, and Hydrogen with H^+ and then add OH^- to each side of the reaction to neutralize H^+ (it's basic conditions now!)

$$5\ OH^{-1} + H_2O + Cr(OH)_3 \longrightarrow CrO_4^{2-} + 3\ e^- + 5\ H^+ + 5\ OH^-$$

$$2\ e^- + Br_2 \longrightarrow 2\ Br^-$$

Now balance the total equation by eliminating the electrons (e^-) (and combine H^+ and OH^- to form water)

This is water!

$$2\ e^- + Br_2 \longrightarrow 2\ Br^-$$

$$5\ OH^{-1} + H_2O + Cr(OH)_3 \longrightarrow CrO_4^{2-} + 3\ e^- + 5\ H^+ + 5\ OH^-$$

Multiply the first reaction by 3
Multiply the second reaction by 2

$$6\ e^- + 3\ Br_2 \longrightarrow 6\ Br^-$$

$$10\ OH^- + 2\ H_2O + 2\ Cr(OH)_3 \longrightarrow 2\ CrO_4^{2-} + 6\ e^- + 10\ H_2O$$

Subtract the differences

$$10\ OH^- + 3\ Br_2 + 2\ Cr(OH)_3 \longrightarrow 2\ CrO_4^{2-} + 6\ Br^- + 8\ H_2O$$