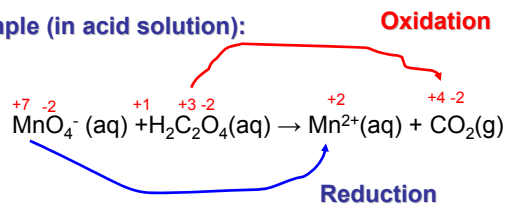


This week's schedule

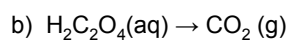
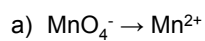
Group 1: Lab Synthesis

Group 2: Tutorial Stoichiometry 1 & 2

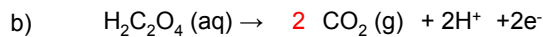
Redox Example (in acid solution):



∴ Half reactions are:



Charge_{LHS} = 8 - 1 = +7 Charge_{RHS} = +2 ∴ Add 5e⁻ to reactant side
(Reduction as expected)



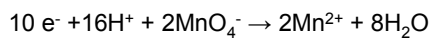
Charge_{LHS} = 0

Charge_{RHS} = +2

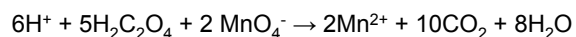
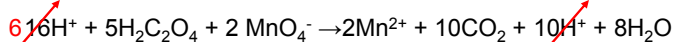
• Products gain 2 e⁻;

• • **(Oxidation as expected)**

c) Multiply reduction reaction by x2 and the oxidation reaction by x5 so that both reactions involve 10 e⁻



+



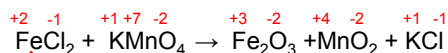
Mass check: 2Mn, 28O, 10C and 16H on both sides

Charge_{LHS} = +6 - 2 = +4

Charge_{RHS} = +4

Example (in basic solution):

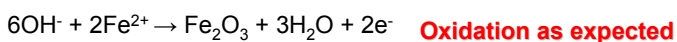
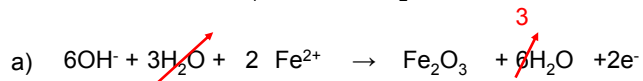
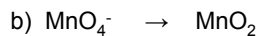
Reduction

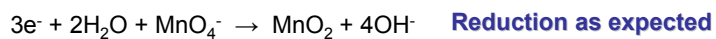
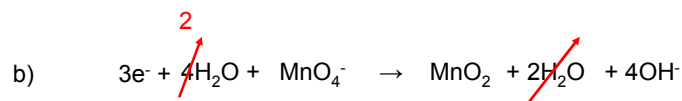


Oxidation

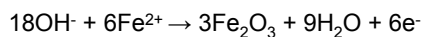
Drop the counter ions whose oxidation state doesn't change: K⁺, Cl⁻.

Write the two half reactions:

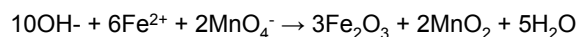
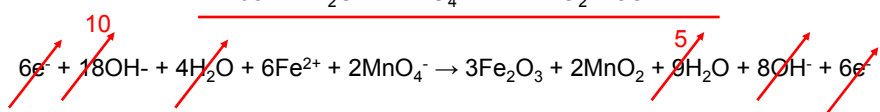
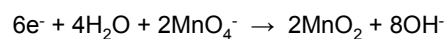




Multiply a) by x3 and b) by x2 to balance the number of electrons



+



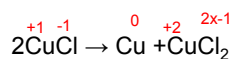
Mass balance: ✓

Charge balance: ✓

Disproportionation Reactions

For elements with 3 or more oxidation states, an element in an intermediate oxidation state can be both oxidized and reduced (a redox reaction known as disproportionation)

Example:



Here oxidation numbers help establish if redox reaction is also a disproportionation reaction

Such reactions can sometimes be difficult to balance

Example: Balance $\text{P}_4 \rightarrow \text{PH}_3 + \text{H}_2\text{PO}_2^-$ in basic solution

Hint: Rewrite the equation with the disproportionating species written as reacting with itself

