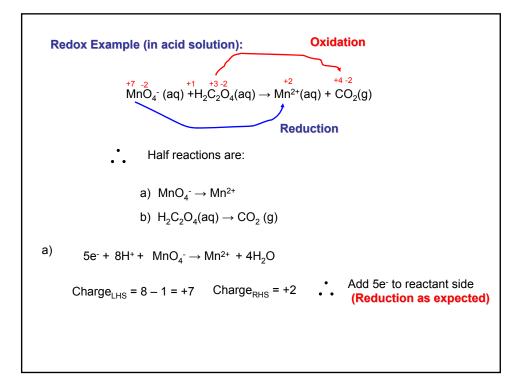
## This week's schedule

Group 1: Lab Synthesis

Group 2: Tutorial Stoichiometry 1 & 2



b) 
$$H_2C_2O_4$$
 (aq)  $\rightarrow$  2  $CO_2$  (g) + 2H<sup>+</sup> +2e<sup>-</sup>

• Products gain 2 e<sup>-</sup>;

Charge<sub>LHS</sub> = 0 Charge<sub>RHS</sub> = +2 • (Oxidization as expected)

c) Multiply reduction reaction by x2 and the oxidation reaction by x 5 so that both reactions involve 10 e-

10 e<sup>-</sup> +16H<sup>+</sup> + 2MnO<sub>4</sub><sup>-</sup> 
$$\rightarrow$$
 2Mn<sup>2+</sup> + 8H<sub>2</sub>O + 
$$5H_2C_2O_4 \rightarrow 10CO_2 + 10H^+ + 10e^-$$

$$616H^{+} + 5H_{2}C_{2}O_{4} + 2 MnO_{4}^{-} \rightarrow 2Mn^{2+} + 10CO_{2} + 10H^{+} + 8H_{2}O$$

$$6H^{+} + 5H_{2}C_{2}O_{4} + 2 MnO_{4}^{-} \rightarrow 2Mn^{2+} + 10CO_{2} + 8H_{2}O$$

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

Charge<sub>LHS</sub> = 
$$+6 - 2 = +4$$
 Charge<sub>RHS</sub> =  $+4$ 

Drop the counter ions whose oxidation state doesn't change: K+, Cl-. Write the two half reactions:

a) Fe
$$^{+2}$$
  $\rightarrow$  Fe $_2$ O $_3$ 

b) 
$$MnO_4^- \rightarrow MnO_2$$

b) 
$$MnO_4^- \rightarrow MnO_2$$
  
a)  $6OH^- + 3H_2O + 2 Fe^{2+} \rightarrow Fe_2O_3 + 9H_2O + 2e^{-1}$ 

$$6 OH^- + 2 Fe^{2+} \rightarrow Fe_2O_3 + 3H_2O + 2e^- \quad \textbf{Oxidation as expected}$$

b) 
$$3e^{-} + \frac{1}{2}H_2O + MnO_4^{-} \rightarrow MnO_2 + 2H_2O + 4OH^{-}$$

$$3e^{-} + 2H_2O + MnO_4^{-} \rightarrow MnO_2 + 4OH^{-}$$
 Reduction as expected

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

$$18OH^{-} + 6Fe^{2+} \rightarrow 3Fe_{2}O_{3} + 9H_{2}O + 6e^{-} + 6e^{-} + 4H_{2}O + 2MnO_{4}^{-} \rightarrow 2MnO_{2} + 8OH^{-}$$

$$\frac{5}{6e^{-} + 18OH - + 4H_{2}O + 6Fe^{2+} + 2MnO_{4}^{-} \rightarrow 3Fe_{2}O_{3} + 2MnO_{2} + 9H_{2}O + 8OH^{-} + 6e^{-}}$$

10OH- + 6Fe
$$^{2+}$$
 + 2MnO $_4$ -  $\rightarrow$  3Fe $_2$ O $_3$  + 2MnO $_2$  + 5H $_2$ O

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 know 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  $P_4 \rightarrow PH_3 + H_2PO_2$  in basic solution

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

Half reactions:

$$12e^{-} + 12H_{2}O + P_{4} \rightarrow 4PH_{3} + 12OH^{-}$$

$$8OH - +8H_2O + P_4 \rightarrow 4 H_2PO_2^- + 8H_2O + 4e^-$$
 (X 3)

12.  
12e<sup>-</sup> +12H<sub>2</sub>O + 4P<sub>4</sub> + 24OH<sup>-</sup> 
$$\rightarrow$$
 4PH<sub>3</sub> + 12H<sub>2</sub>PO<sub>2</sub><sup>-</sup> +12OH<sup>-</sup> + 12e<sup>-</sup>  
= 12H<sub>2</sub>O + 4P<sub>4</sub> +12OH<sup>-</sup>  $\rightarrow$  4PH<sub>3</sub> +12H<sub>2</sub>PO<sub>2</sub><sup>-</sup>

Divide equation by 4

$$3H_2O + P_4 + 3OH^- \rightarrow PH_3 + 3H_2PO_2^-$$

Mass check: 9H, 6O and 4 P on both sides of the equation

$$Charge_{LHS} = -3$$

$$Charge_{RHS} = -3$$
Balanced