Non Sibi High School

Andover's Chem 550/580: Advanced Chemistry

Chapter 16, Review Quiz 1 Answers

1

Determine all oxidation numbers in:

a. As^{3-}
b. F_2
c. $HO_2{}^-$
d. $IO_4{}^-$
e. KH

a. As = -3
b. F = 0
c. H = +1, O = -1
d. O = -2, I = +7
e. K = +1, H = -1

2

Determine all oxidation numbers, identify the reducing agent and oxidizing agent, and balance the following equation that occurs in aqueous acidic solution using the smallest possible whole-number coefficients:

$$UO^{2+} + MnO_4^- \longrightarrow UO_2{}^{2+} + MnO_2$$

$$UO^{2+}(O = -2, U = +4) + MnO_4^-(O = -2, Mn = +7) \longrightarrow$$
$$UO_2{}^{2+}(O = -2, U = +6) + MnO_2(O = -2, Mn = +4)$$

reducing agent = UO^{2+} (U increases from +4 to +6)
oxidizing agent = $MnO_4{}^-$ (Mn decreases from +7 to +4)

oxidation: $(UO^{2+} + H_2O \longrightarrow UO_2{}^{2+} + 2H^+ + 2e^-) \times 3$
reduction: $(MnO_4{}^- + 4H^+ + 3e^- \longrightarrow MnO_2 + 2H_2O) \times 2$
balanced equation: $3UO^{2+} + 2MnO_4^- + 2H^+ \longrightarrow 3UO_2{}^{2+} + 2MnO_2 + H_2O$

3

Rank the solid alkali metals K, Li, and Na from weakest to strongest reducing agent under standard conditions. Justify your answer using a table of standard reduction potentials.

weakest = Na(+2.71 V) < K(+2.92 V) < Li (+3.05 V) = strongest

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Rank the aqueous cations Ag^{+}, Al^{3+}, and Cd^{2+} from weakest to strongest oxidizing agent under standard conditions. Justify your answer using a table of standard reduction potentials.

weakest = Al^{3+}(-1.66 V) < Cd^{2+}(-0.40 V) < Ag^{+}(+0.80 V) = strongest

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For each spontaneous reaction below, calculate E°_{cell} and then balance the equation.

a. chlorine gas + aqueous potassium bromide
b. solid aluminum metal + aqueous hydrochloric acid
c. solid gold metal + aqueous nitric acid
d. solid zinc metal + aqueous cadmium(II) nitrate

a. $E^{\circ}_{cell} = (+1.36\,V) + (-1.07\,V) = 0.29\,V$

$$Cl_2(g) + 2Br^-(aq) \longrightarrow 2Cl^-(aq) + Br_2(l)$$

b. $E^{\circ}_{cell} = (+1.66\,V) + (0\,V) = 1.66\,V$

$$2Al(s) + 6H^+(aq) \longrightarrow 2Al^{3+}(aq) + 3H_2(g)$$

c. nonspontaneous
d. $E^{\circ}_{cell} = (+0.76\,V) + (-0.40\,V) = 0.36\,V$

$$Zn(s) + Cd^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Cd(s)$$

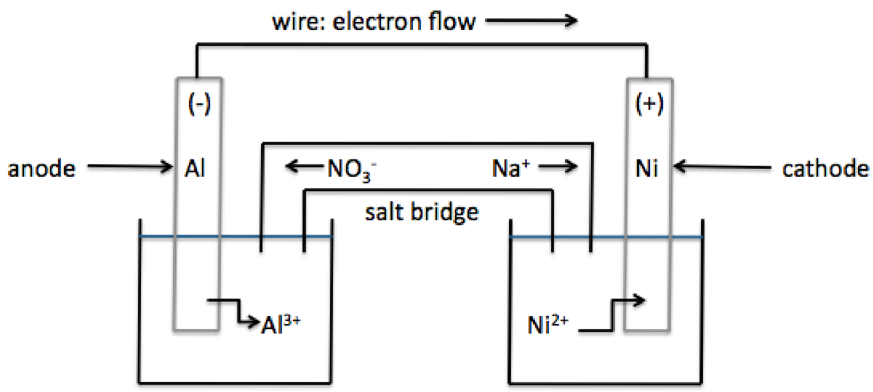
6

A galvanic cell was constructed using a strip of nickel metal and a strip of aluminum metal, a 1 M solution of $NiSO_4$ and a 1 M solution of $Al(NO_3)_3$, and an aqueous solution of $NaNO_3$ in the salt bridge. For the spontaneous reaction that occurred, calculate E°_{cell} and ΔG°, then balance the equation. Also sketch the galvanic cell.

$$E^\circ_{cell} = (+1.66\,V) + (-0.25\,V) = 1.41\,V$$

$$\Delta G^\circ = -6\left(\frac{96,500\,C}{1\,mol}\right)\left(\frac{1.41\,J}{1\,C}\right) = -8.16 \times 10^5\,J/mol = -816\,kJ/mol$$

$$2Al(s) + 3Ni^{2+}(aq) \longrightarrow 2Al^{3+}(aq) + 3Ni(s)$$



7

Calculate the minimum voltage required to bring about the reaction $Cr^{3+}(aq) + Ag(s) \longrightarrow Cr(s) + Ag^{+}(aq)$ by electrolysis under standard conditions, then balance the equation.

$$E^\circ_{cell} = (-0.74\,V) + (-0.80\,V) = -1.54\,V$$

$$\text{minimum voltage} = |E^\circ_{cell}| = 1.54\,V$$

$$Cr^{3+}(aq) + 3Ag(s) \longrightarrow Cr(s) + 3Ag^{+}(aq)$$

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For the electrolysis of molten NaI, write the half-reaction that occurs at the anode and the half-reaction that occurs at the cathode, then balance the equation.

oxidation (at anode): $(2I^- \longrightarrow I_2 + 2e^-) \times 1$
reduction (at cathode): $(Na^+ + 1e^- \longrightarrow Na) \times 2$
balanced equation: $2I^- + 2Na^+ \longrightarrow I_2 + 2Na$

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a. For the electrolysis of aqueous NaI, write the half-reaction that occurs at the anode and the half-reaction that occurs at the cathode, then balance the equation.

b. If the electrolysis proceeds for 2.65 days using a current of 3.75 A, how many kilograms of I_2 will be produced?

a. $Na^+(aq)$ = electrolysis spectator ion, E°_{ox} = -0.53 V for the oxidation of $I^-(aq)$ is less negative than E°_{ox} for the oxidation of water:

oxidation (at anode): $2I^-(aq) \longrightarrow I_2(s) + 2e^-$
reduction (at cathode): $2H_2O(l) + 2e^- \longrightarrow H_2(g) + 2OH^-(aq)$
balanced equation: $2I^-(aq) + 2H_2O(l) \longrightarrow I_2(s) + H_2(g) + 2OH^-(aq)$

b. oxidation (at anode): $2I^-(aq) \longrightarrow I_2(s) + 2e^-$

$$2.65\,\mathrm{d}\left(\frac{24\,\mathrm{h}}{1\,\mathrm{d}}\right)\left(\frac{60\,\mathrm{min}}{1\,\mathrm{h}}\right)\left(\frac{60\,\mathrm{s}}{1\,\mathrm{min}}\right)\left(\frac{3.75\,\mathrm{C}}{1\,\mathrm{s}}\right)\left(\frac{1\,\mathrm{mol\,e^-}}{96,500\,\mathrm{C}}\right)\left(\frac{1\,\mathrm{mol\,I_2}}{2\,\mathrm{mol\,e^-}}\right)\left(\frac{253.8\,\mathrm{g}}{1\,\mathrm{mol\,I_2}}\right)\left(\frac{1\,\mathrm{kg}}{1000\,\mathrm{g}}\right) = 1.13\,\mathrm{kg}$$

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a. For the electrolysis of aqueous $LiNO_3$, write the half-reaction that occurs at the anode and the half-reaction that occurs at the cathode, then balance the equation.

b. Using a current of 0.285 A, how many hours must the electrolysis proceed to produce 444 mL of oxygen gas, measured at 28°C and 724 torr?

c. If 40.0 milligrams of hydrogen gas are produced when the electrolysis proceeds for 18.5 minutes, calculate the current used in the electrolysis.

a. $Li^+(aq)$ and $NO_3^-(aq)$ = electrolysis spectator ions:

oxidation (at anode): $2H_2O(l) \longrightarrow O_2(g) + 4H^+(aq) + 4e^-$
reduction (at cathode): $(2H_2O(l) + 2e^- \longrightarrow H_2(g) + 2OH^-(aq)) \times 2$
balanced equation: $2H_2O(l) \longrightarrow 2H_2(g) + O_2(g)$

b. oxidation (at anode): $2H_2O(l) \longrightarrow O_2(g) + 4H^+(aq) + 4e^-$

$$n = \frac{\frac{724}{760}\,\mathrm{atm} \times \frac{444}{1000}\,\mathrm{L}}{0.0821\,\frac{\mathrm{L\cdot atm}}{\mathrm{mol\cdot K}} \times (28 + 273)\,\mathrm{K}} = 0.01712\,\mathrm{mol\,O_2}$$

$$0.01712\,\mathrm{mol\,O_2}\left(\frac{4\,\mathrm{mol\,e^-}}{1\,\mathrm{mol\,O_2}}\right)\left(\frac{96,500\,\mathrm{C}}{1\,\mathrm{mol\,e^-}}\right)\left(\frac{1\,\mathrm{s}}{0.285\,\mathrm{C}}\right)\left(\frac{1\,\mathrm{min}}{60\,\mathrm{s}}\right)\left(\frac{1\,\mathrm{h}}{60\,\mathrm{min}}\right) = 6.44\,\mathrm{h}$$

c. reduction (at cathode): $2H_2O(l) + 2e^- \longrightarrow H_2(g) + 2OH^-(aq)$

$$18.5\,\text{min}\left(\frac{60\,\text{s}}{1\,\text{min}}\right) = 1110\,\text{s}$$

$$40.0\,\text{mg}\left(\frac{1\,\text{g}}{1000\,\text{mg}}\right)\left(\frac{1\,\text{mol}\,H_2}{2.016\,\text{g}}\right)\left(\frac{2\,\text{mol}\,e^-}{1\,\text{mol}\,H_2}\right)\left(\frac{96,500\,\text{C}}{1\,\text{mol}\,e^-}\right) = 3829\,\text{C}$$

$$\text{current} = \frac{3829\,\text{C}}{1110\,\text{s}} = 3.45\,\text{C/s} = 3.45\,\text{A}$$

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Determine all oxidation numbers, identify the reducing agent and oxidizing agent, and balance the following equation that occurs in aqueous basic solution using the smallest possible whole-number coefficients:

$$IO^- + Cr(OH)_3 \longrightarrow I_2 + CrO_4^{2-}$$

$$IO^-(O = -2, I = +1) + Cr(OH)_3(H = +1, O = -2, Cr = +3) \longrightarrow$$
$$I_2(I = 0) + CrO_4^{2-}(O = -2, Cr = +6)$$

reducing agent = $Cr(OH)_3$ (U increases from +3 to +6)
oxidizing agent = IO^- (I decreases from +1 to 0)

oxidation: $(Cr(OH)_3 + H_2O \longrightarrow CrO_4^{2-} + 5H^+ + 3e^-) \times 2$
reduction: $(IO^- + 4H^+ + 2e^- \longrightarrow I_2 + 2H_2O) \times 3$
balanced equation (acidic): $2Cr(OH)_3 + 6IO^- + 2H^+ \longrightarrow 2CrO_4^{2-} + 3I_2 + 4H_2O$

$$2Cr(OH)_3 + 6IO^- + 2H^+ + 2OH^- \longrightarrow 2CrO_4^{2-} + 3I_2 + 4H_2O + 2OH^-$$

$$2Cr(OH)_3 + 6IO^- + 2H_2O \longrightarrow 2CrO_4^{2-} + 3I_2 + 4H_2O + 2OH^-$$
$$2Cr(OH)_3 + 6IO^- \longrightarrow 2CrO_4^{2-} + 3I_2 + 2H_2O + 2OH^-$$

12

Calculate K_c for the reaction $Fe^{2+}(aq) + Cd(s) \longrightarrow Cd^{2+}(aq) + Fe(s)$ at 25°C.

$$E^\circ_{cell} = (-0.44\,\text{V}) + 0.40\,\text{V} = -0.04\,\text{V}$$
$$K_{eq} = K_c:$$
$$-0.04\,\text{V} = \frac{0.0257\,\text{V}}{2}\ln K_c$$
$$K_c = 0.04$$

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A galvanic cell is constructed at 25°C that utilizes the reaction $2Fe^{3+}(aq) + Cu(s) \longrightarrow 2Fe^{2+}(aq) + Cu^{2+}(aq)$.

a. If all aqueous ions in the cell have an initial concentration of 1 M, calculate E°_{cell}.

b. Predict whether E_{cell} will increase, decrease, or remain unchanged if the initial concentration of Fe^{3+} is 1 M, but the initial concentrations of Fe^{2+} and Cu^{2+} are both changed to 0.075 M.

c. Calculate E_{cell} if the initial concentrations of all aqueous ions are 0.025 M.

a. $E^{\circ}_{cell} = 0.77\,V + (-0.34\,V) = 0.43\,V$

b. decrease product concentrations only = Q_c decreases = E_{cell} increases

c.

$$E_{cell} = 0.43\,V - \frac{0.0257\,V}{2} \ln \frac{(0.025)^2(0.025)}{(0.025)^2} = 0.48\,V$$