Periodic Table of the Elements

1A 1	2A 2	3B 3	4B 4	5B 5	6B 6	7B 7	8B 8	8B 9	8B 10	1B 11	2B 12	3A 13	4A 14	5A 15	6A 16	7A 17	8A 18
1 H 1.0079																	2 He 4.0026
3 Li 6.941	4 Be 9.0122											5 B 10.811	6 C 12.011	7 N 14.0067	8 O 15.9994	9 F 18.9984	10 Ne 20.1797
11 Na 22.9898	12 Mg 24.3050											13 Al 26.9815	14 Si 28.0855	15 P 30.9738	16 S 32.066	17 Cl 35.4527	18 Ar 39.948
19 K 39.0983	20 Ca 40.078	21 Sc 44.9559	22 Ti 47.88	23 V 50.9415	24 Cr 51.9961	25 Mn 54.9380	26 Fe 55.847	27 Co 58.9332	28 Ni 58.69	29 Cu 63.546	30 Zn 65.39	31 Ga 69.723	32 Ge 72.61	33 As 74.9216	34 Se 78.96	35 Br 79.904	36 Kr 83.80
37 Rb 85.4678	38 Sr 87.62	39 Y 88.9059	40 Zr 91.224	41 Nb 92.9064	42 Mo 95.94	43 Tc (98)	44 Ru 101.07	45 Rh 102.9055	46 Pd 106.42	47 Ag 107.8682	48 Cd 112.411	49 In 114.82	50 Sn 118.710	51 Sb 121.75	52 Te 127.60	53 I 126.9045	54 Xe 131.39
55 Cs 132.9054	56 Ba 137.327	57 La 138.9055	72 Hf 178.49	73 Ta 180.9479	74 W 183.85	75 Re 186.207	76 Os 190.2	77 Ir 192.22	78 Pt 195.08	79 Au 196.9665	80 Hg 200.59	81 Tl 204.3833	82 Pb 207.2	83 Bi 208.9804	84 Po (209)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra (226)	89 Ac (227)	104 Rf (261)	105 Db (262)	106 Sg (263)	107 Bh (262)	108 Hs (265)	109 Mt (266)									

58 Ce 140.115	59 Pr 140.9076	60 Nd 144.24	61 Pm (145)	62 Sm 150.36	63 Eu 151.965	64 Gd 157.25	65 Tb 158.9253	66 Dy 162.50	67 Ho 164.9303	68 Er 167.26	69 Tm 168.9342	70 Yb 173.04	71 Lu 174.967
90 Th 232.0381	91 Pa 231.0359	92 U 238.0289	93 Np (237)	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (260)

This print-out should have 8 questions. Multiple-choice questions may continue on the next column or page – find all choices before answering. The due time is Central time.

Std emf 12 17
20:07, basic, multiple choice, < 1 min, wording-variable.
001

Predict the standard emf of the given cell
$C(gr) \,|\, Sn^{4+}(aq), Sn^{2+}(aq) \,||\, Pb^{4+}(aq), Pb^{2+} \,|\, Pt(s)$

1. +1.52 V **correct**

2. +0.75 V

3. +0.37 V

4. +0.52 V

5. +2.14 V

6. +1.34 V

Explanation:
Identify the cathode (right-side) and anode (left-side) reactions and potentials from the cell diagram.
At the cathode,
$Pb^{4+}(aq) + 2\,e^- \rightarrow Pb^{2+}(aq) \quad E^\circ = +1.67$ V
At the anode,
$Sn^{2+}(aq) \rightarrow Sn^{4+}(aq) + 2\,e^- \; -E^\circ = -0.15$ V

$$\begin{aligned} E^\circ_{\text{cell}} &= E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} \\ &= +1.67\text{ V} - (+0.15\text{ V}) \\ &= +1.52\text{ V} \end{aligned}$$

Msci 21 1220
20:07, general, multiple choice, > 1 min, fixed.
002
Consider the half-reactions
$Mn^{2+} + 2\,e^- \rightarrow Mn \quad E^0 = -1.029$ V
$Ga^{3+} + 3\,e^- \rightarrow Ga \quad E^0 = -0.560$ V
$Fe^{2+} + 2\,e^- \rightarrow Fe \quad E^0 = -0.409$ V
$Sn^{2+} + 2\,e^- \rightarrow Sn \quad E^0 = -0.136$ V

Of the species listed, the strongest oxidizing agent is

1. Sn^{2+} **correct**

2. Mn^{2+}

3. Mn

4. Sn

5. Ga^{+3}

Explanation:
Oxidizing agents get reduced. As E_0 increases, the easier it is for the species to be reduced. Since Sn^{2+} has the biggest E_0, it is reduced the easiest, making it the strongest oxidizing agent.

ChemPrin3e T12 36
20:09, basic, multiple choice, < 1 min, fixed.
003
Consider the cell

$$Zn(s) \,|\, Zn^{2+}(aq) \,||\, Fe^{2+}(aq) \,|\, Fe(s)$$

at standard conditions.
Calculate the value of ΔG°_r for the reaction that occurs when current is drawn from this cell.

1. $-\,62$ kJ · mol^{-1} **correct**

2. $-\,230$ kJ · mol^{-1}

3. $+\,62$ kJ · mol^{-1}

4. $+\,230$ kJ · mol^{-1}

5. $-\,31$ kJ · mol^{-1}

Explanation:

Mlib 08 0085
20:12, basic, multiple choice, > 1 min, fixed.
004
A battery has two terminals labeled positive and negative.

As the battery discharges, electrons flow from the __?__ terminal to the __?__ terminal through the external circuit and __?__ reaction occurs at the positive terminal.

1. positive; negative; a reduction

2. postive; negative; an oxidation

3. negative; positive; a reduction **correct**

4. negative; positive; an oxidation

5. positive; negative; an acid/base

Explanation:
In a voltaic cell electrons flow from the negative to the positive terminals. Reduction occurs at the positive terminal.

Mlib 08 0097
20:11, general, multiple choice, > 1 min, fixed.
005
Which of the following batteries could not be recharged?

1. dry cell **correct**

2. lead storage battery

3. nickel-cadium battery

Explanation:

Msci 21 0002
20:08, general, multiple choice, > 1 min, fixed.
006
Calculate the potential for the cell indicated:
Fe | Fe^{2+} (10^{-3} M) || Pb^{2+} (10^{-5} M) | Pb

$Pb^{2+} + 2\,e^- \rightarrow Pb \qquad E^0 = -0.126$ V

$Fe^{2+} + 2\,e^- \rightarrow Fe \qquad E^0 = -0.440$ V

1. 0.255 V **correct**

2. 0.432 V

3. 0.373 V

4. 0.196 V

5. 0.284 V

Explanation:
The overall reaction is

$$Fe + Pb^{2+} \rightarrow Fe^{2+} + Pb$$

Please notice that since the concentrations are not 1 M, the Nernst equation must be used.
In this cell notation, the anode is located on the left of the salt bridge || and the cathode on the right. So first calculate

$$\begin{aligned} E^0_{\text{cell}} &= E_{\text{cathode}} - E^0_{\text{anode}} \\ &= -0.126 \text{ V} - (-0.440) \text{ V} = 0.314 \text{ V} \end{aligned}$$

Using the Nernst Equation

$$\begin{aligned} E_{\text{cell}} &= E^0_{\text{cell}} - \frac{0.05916}{n} \log Q \\ &= 0.314 \text{ V} - \frac{0.05916}{2} \log\left(\frac{[Fe^{2+}]}{[Pb^{2+}]}\right) \\ &= 0.314 \text{ V} - \frac{0.05916}{2} \log\left(\frac{10^{-3}}{10^{-5}}\right) \\ &= 0.25484 \text{ V} \end{aligned}$$

Msci 21 0606
20:05, general, multiple choice, > 1 min, fixed.
007
What weight of Cl_2 gas will be produced by electrolysis of molten NaCl when a current of 4.35 amps flows through it for 15.0 hours? (Cl = 35.457 g/mol)

1. 86.3 g **correct**

2. 19.8 g

3. 0.0250 g

4. 1.44 g

5. 43.2 g

Explanation:

Msci 04 0900
20:01, general, multiple choice, > 1 min, fixed.
008

Using the smallest possible integer coefficients to balance the redox equation

$$MnO_4^- + C_2O_4^{-2} \to Mn^{+2} + CO_2$$

(acidic solution), the coefficient for $C_2O_4^{2-}$ is

1. 5. **correct**

2. 2.

3. 4.

4. 7.

5. The correct coefficient is not given.

Explanation:
The oxidation number of C changes from +3 to +4, so C is oxidized. The oxidation number of Mn changes from +7 to +2, so Mn is reduced. We set up oxidation and reduction half-reactions:

Red: $MnO_4^- \to Mn^{2+}$

Oxid: $C_2O_4^{2-} \to CO_2$

Mn atoms are balanced. We need 2 CO_2 molecules to balance C:

Oxid: $C_2O_4^{2-} \to 2\,CO_2$

Since this is an acidic solution, we use H_2O and H^+ to balance O and H atoms, adding the H_2O to the side needing oxygen:

Red: $8\,H^+ + MnO_4^- \to Mn^{2+} + 4\,H_2O$

Oxid: $C_2O_4^{2-} \to 2\,CO_2$

We balance the total charge in each half-reaction by adding electrons. In the preceding reduction reaction there is a total charge of +7 on the left and +2 on the right. Five electrons are added to the left:

Red: $5\,e^- + 8\,H^+ + MnO_4^- \to Mn^{2+} + 4\,H_2O$

Oxid: $C_2O_4^{2-} \to 2\,CO_2 + 2\,e^-$

The number of electrons gained by Mn must equal the number of electrons lost by C. We multiply the reduction reaction by 2 and the oxidation reaction by 5 to balance the electrons:

Red: $10\,e^- + 16\,H^+ + 2\,MnO_4^- \to 2\,Mn^{2+} + 8\,H_2O$

Oxid: $5\,C_2O_4^{2-} \to 10\,CO_2 + 10\,e^-$

Adding the half-reactions gives the overall balanced equation:

$$5\,C_2O_4^{2-} + 16\,H^+ + 2\,MnO_4^- \to 10\,CO_2 + 2\,Mn^{2+} + 8\,H_2O$$