Teacher's Tools® Chemistry
Electrochemistry: Voltaic Cells: Worksheet 3

KEY

B 1. If a reactant loses electrons it is
(A) reduced.
(B) oxidized.
(C) disproportionated.
(D) the cathode.
(E) a catalyst.

E 2. The anode is where
(A) electrolyte concentrations are lowest.
(B) electrolyte concentrations are highest.
(C) the Nernst potential is measured.
(D) anions are produced.
(E) oxidation occurs.

C 3. All of the following statements concerning voltaic cells are true EXCEPT
(A) a voltaic cell consists of two half-cells.
(B) a voltaic cell can serve as an energy source. (source of electrical energy)
(C) the electrodes in a voltaic cell must be composed of dissimilar metals.
(D) the two half-cells in a voltaic cell are connected by a salt bridge. → not necessarily
(E) reduction occurs at the cathode in a voltaic cell.

D 4. Write a balanced chemical equation for the overall reaction represented below.

$$Al \mid Al^{3+} \parallel Zn \mid Zn^{2+}$$

$2(Al \rightarrow Al^{3+} + 3e^-)$

$3(Zn^{2+} + 2e^- \rightarrow Zn)$

(A) $3Zn(s) + 2Al(s) \rightarrow 2Al^{+}(aq) + 3Zn^{2+}(aq)$
(B) $Zn(s) + 2Al^{+}(aq) \rightarrow 2Al(s) + Zn^{2+}(aq)$
(C) $3Zn^{2+}(aq) + 2Al^{3+}(aq) \rightarrow 2Al(s) + 3Zn(s)$
(D) $3Zn^{2+}(aq) + 2Al(s) \rightarrow 2Al^{3+}(aq) + 3Zn(s)$
(E) $Zn(s) + Al^{3+}(aq) \rightarrow Al(s) + Zn^{2+}(aq)$

C WM 5. Write a balanced chemical equation for the overall reaction represented below.

$$Pt \mid Fe^{2+}, Fe^{3+} \parallel Cl^- \mid AgCl \mid Ag$$

ox $Fe^{2+} \rightarrow Fe^{3+} + e^-$

$AgCl \rightarrow Ag^{+} + Cl^-$

reduction of Ag^+

(A) $AgCl(s) + Fe^{3+}(aq) + 2e^- \rightarrow Ag(s) + Fe^{2+}(aq) + Cl^-(aq)$
(B) $AgCl(s) + Fe^{3+}(aq) \rightarrow Ag(s) + Fe^{2+}(aq) + Cl^-(aq)$
(C) $AgCl(s) + Fe^{2+}(aq) \rightarrow Ag(s) + Fe^{3+}(aq) + Cl^-(aq)$
(D) $Ag(s) + Fe^{3+}(aq) + Cl^-(aq) \rightarrow AgCl(s) + Fe^{2+}(aq)$
(E) $Ag(s) + Fe^{2+}(aq) + Cl^-(aq) \rightarrow AgCl(s) + Fe^{3+}(aq)$

C 6. What is the correct cell notation for the reaction below?

$$Cd^{2+}(aq) + Ni(s) \rightarrow Cd(s) + Ni^{2+}(aq)$$

(A) $Cd^{2+} \mid Cd \parallel Ni \mid Ni^{2+}$
(B) $Cd \mid Cd^{2+} \parallel Ni^{2+} \mid Ni$
(C) $Ni \mid Ni^{2+} \parallel Cd^{2+} \mid Cd$
(D) $Ni \mid Cd^{2+} \parallel Ni^{2+} \mid Cd$
(E) $Cd^{2+} \mid Cd \parallel Ni \mid Ni^{2+}$

A 7. Which species is the best oxidizing agent? (*refer to a voltage table*)
(A) bromine Br_2
(B) sodium ion Na^+
(C) potassium metal K
(D) fluoride ion F^-
(E) carbon dioxide CO_2

D 8. Which species is the best reducing agent? (*refer to a voltage table*)
(A) sodium ion Na^+
(B) sulfide ion S^{2-}
(C) chlorine Cl_2
(D) sodium metal Na
(E) silver metal Ag

B 9. The more _____ E°_{red} is for a species, the stronger the oxidizing strength of the species.
(A) voltaic
(B) positive
(C) dilute
(D) negative
(E) concentrated

C 10. Which of the followings species are oxidizing agents: I_2, F^-, $Cr_2O_7^{2-}$, Ni, and Li?
(A) I_2 only
(B) I_2 and F^-
(C) I_2 and $Cr_2O_7^{2-}$
(D) F^-, Ni, and Li
(E) Ni and Li

A 11. Which of the following species are reducing agents: K, H^+, MnO_4^-, Cl_2, Sn^{4+}?
(A) K only
(B) K and Cl_2
(C) H^+ and Sn^{4+}
(D) MnO_4^- and Cl_2
(E) Cl_2 and Sn^{4+}

D 12. Using the reduction half-reactions, write an overall reaction in the direction that is spontaneous. Calculate $E°$.

$Ag + I^- \rightarrow AgI$

$AgI(s) + e^- \rightarrow Ag(s) + I^-(aq)$ $\quad E^*_{red} = -0.152$ V

$Co^{3+}(aq) + e^- \rightarrow Co^{2+}(aq)$ $\quad E^*_{red} = +1.953$ V

$E^*_{ox} = +0.152$

1.953
.152
2.105

(A) $Co^{2+}(aq) + AgI(s) \rightarrow Co^{3+}(aq) + Ag(s) + I^-(aq)$ $\quad E° = -2.105$ V
(B) $Co^{2+}(aq) + Ag(s) + I^-(aq) \rightarrow Co^{3+}(aq) + AgI(s)$ $\quad E° = -1.801$ V
(C) $Co^{3+}(aq) + Ag(s) + I^-(aq) \rightarrow Co^{2+}(aq) + AgI(s)$ $\quad E° = +1.801$ V
(D) $Co^{3+}(aq) + Ag(s) + I^-(aq) \rightarrow Co^{2+}(aq) + AgI(s)$ $\quad E° = +2.105$ V
(E) $Co^{3+}(aq) + AgI(s) \rightarrow Co^{2+}(aq) + Ag(s) + I^-(aq)$ $\quad E° = -2.105$ V

E 13. The cell voltage, $E°$, is _____ for a reaction taking place in a voltaic cell.

(A) reductive
(B) oxidative
(C) negative
(D) zero
(E) positive

spontaneous

E 14. Use the standard reduction potentials below to determine which of the following metals can be oxidized by 1 M HCl: Zn, Sn, Cu, and ~~Fe~~ Al

$2H^+ + 2e^- \rightarrow H_2$ $\quad E° = 0.00$ (reduct)

$Al^{3+}(aq) + 3e^- \rightarrow Al(s)$ $\quad E^*_{red} = -1.68$ V
$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$ $\quad E^*_{ox} = -0.762$ V
$Sn^{2+}(aq) + 2e^- \rightarrow Sn(s)$ $\quad E^*_{ox} = -0.141$ V
$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ $\quad E^*_{red} = +0.339$ V

(A) Al only
(B) Cu only
(C) Al and Cu
(D) Zn and Sn
(E) Al, Zn, and Sn

A 15. The standard reduction potentials for $F_2(g)$ and $Br_2(l)$ are +2.889 V and 1.077 V, respectively. Write the chemical equation and calculate $E°$ for a voltaic cell based on fluorine and bromine.

(A) $F_2(g) + 2Br^-(aq) \rightarrow 2F^-(aq) + Br_2(g)$ $\quad E° = +1.812$ V
(B) $F_2(g) + Br_2(g) \rightarrow 2F^-(aq) + 2Br^-(aq)$ $\quad E° = -1.812$ V
(C) $F_2(g) + Br_2(g) \rightarrow 2F^-(aq) + 2Br^-(aq)$ $\quad E° = +1.812$ V
(D) $2F^-(aq) + 2Br^-(aq) \rightarrow F_2(g) + Br_2(g)$ $\quad E° = +1.812$ V
(E) $Br_2(g) + 2F^-(aq) \rightarrow 2Br^-(aq) + F_2(g)$ $\quad E° = -1.812$ V

$F_2 + 2e^- \rightarrow 2F^-$ $\quad +2.889$
$Br_2 + 2e^- \rightarrow 2Br^-$ $\quad +1.079$

ox $2Br^- \rightarrow 2e^- + Br_2$ $\quad -1.079$

C 16. For a spontaneous reaction,

(A) $\Delta G° > 0$ and $E° > 0$.
(B) $\Delta G° > 0$ and $E° < 0$.
(C) $\Delta G° < 0$ and $E° > 0$.
(D) $\Delta G° < 0$ and $E° < 0$.
(E) $\Delta G° = 0$ and $E° = 0$.

C 17. What is the potential in volts for the spontaneous reaction between the Ag|Ag+ and Zn|Zn^{2+} half-cells?

$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$ $E° = -0.763$ V
$Ag^{+}(aq) + e^- \rightarrow Ag(s)$ $E° = 0.799$ V

(A) -2.361 V
(B) -1.562 V
(C) 1.562 V
(D) 2.361 V

B 18. What is the $[Cu^{2+}]$ in the cell Zn|Zn2+ (0.05M) || Cu2+ (? M) | Cu if the cell voltage is 1.03V?

$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$ $E° = -0.763$ V
$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ $E° = 0.337$ V

(A) 0.12 M
(B) 0.0002 M
(C) 0.05 M
(D) 0.0035 M

C B 19. Given the standard electrode potentials:

$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ $E° = 0.337$ V
$Al^{3+}(aq) + 3e^- \rightarrow Al(s)$ $E° = -1.66$ V

The voltage of the cell: $Al_2(SO_4)_3$ (0.1 M) | Al || Cu | $CuSO_4$ (0.5M) would be

(A) 0.00 V
(B) 1.98 V
(C) 2.01 V
(D) 5.70 V

$Cr_2O_7^{2-}(aq) + H_2S(g) + H^+(aq) \rightarrow Cr^{3+}(aq) + S(s) + H_2O(l)$

X 20. When the equation above is correctly balanced and all coefficients are reduced to their lowest whole-number terms, the coefficient for $H_2O(g)$ is

(A) 2 (B) 4 (C) 6 (D) 8 (E) 14

4(a) ox: $Zn \rightarrow Zn^{2+} + 2e^-$ $E^0 = +0.76$

red $Ag^+ + e^- \rightarrow Ag$ $E^0 = 0.80$

$= +1.56$ volts

(b) $E = E^0 - \frac{0.0592}{n} \log \frac{[Zn^{2+}]}{[Ag^+]^2}$

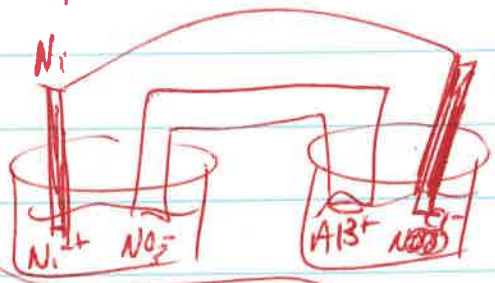
$1.04 = 1.56 - \frac{0.0592}{2} \log\left(\frac{1.00}{x^2}\right)$

$\frac{0.52}{0.0296} = \log \frac{1}{x^2} \Rightarrow \frac{1}{x^2} = 3.7 \times 10^{+17}$

$x = [Ag^+] = 1.64 \times 10^{-9}$ M

(c) $K_{sp} = [Ag^+][Cl^-] = (1.64 \times 10^{-9})(0.100) = 1.64 \times 10^{-10}$

5.



anode
oxidation
$Al \rightarrow Al^{3+} + 3e^-$ $E^0_{ox} = +1.66$

cathode
reduction $Ni^{2+} + 2e^- \rightarrow Ni$ $E^0_{red} = -0.25$ volts

(b) $2Al + 3Ni^{2+} \rightarrow 3Ni + 2Al^{3+}$ $E^0_{cell} = 1.41$ volts

(c) $E = E^0 - \frac{0.0592}{n} \log \frac{[Al^{3+}]^2}{[Ni^{2+}]^3}$

$E = 1.41 - \frac{0.0592}{6} \log\left(\frac{.750^2}{.500^3}\right) = 1.40$

not much of a change after all that math!!

6. (A) $M + Cu^{2+} \rightarrow M^{2+} + Cu \quad E^\circ_{cell} = 1.101$ volts

red

$Cu^{2+} + 2e^- \rightarrow Cu \quad E^\circ_{red} = +0.34$

$M \rightarrow M^{2+} + 2e^- \quad E^\circ_{ox} = ?$

$0.34 + E^\circ_{ox} = 1.101$

$E^\circ_{ox} = 0.761$

SO $M^{2+} + 2e^- \rightarrow M \quad E^\circ_{red} = -0.761$ volts

(B) $M + Cu^{2+} \rightarrow M^{2+} + Cu$

I	mass	1.00	1.00	mass
c	−	−0.80	+0.80	+
E		0.20	1.80	

$$E = E^\circ - \frac{0.0592}{n} \log \frac{[M^{2+}]}{[Cu^{2+}]}$$

$$E = 1.101 - \frac{0.0592}{2} \log\left(\frac{1.80}{0.20}\right)$$

$E = 1.07$ volts

As cell operates,
concentration of reaction ions ↓
concentration of product ions ↑
cell voltage ↓