Honors Chemistry

Answers to Chapter 18 Study Questions

1. a) -3; [$x + 3(1) = 0$] b) 0 c) +4; [$x + 2(-2) = 0$] d) +5; [$x + 3(-2) = -1$]

2. a) Br_2 b) I c) Br d) Br_2 e) I^-

3. First determine whether substances may be oxidized or reduced, then look up E^o_{ox} for reducing agents and E^o_{red} for oxidizing agents and list in order.

Oxidizing agents (are reduced): H^+ ($E^o_{red} = 0$ V).

Reducing agents (are oxidized): Ni ($E^o_{ox} = +0.23$ V), Au ($E^o_{ox} = -1.50$ V), Mg ($E^o_{ox} = +2.37$ V).

Both: Sn^{2+} ($E^o_{red} = -0.14$ V; $E^o_{ox} = -0.15$ V), Fe^{2+} ($E^o_{red} = -0.44$ V, $E^o_{ox} = -0.77$ V), Cl_2 ($E^o_{red} = 1.36$ V, $E^o_{ox} = -1.47$ V,).

Oxidizing agents (list in order of decreasing E^o_{red}): $Cl_2 > H^+ > Sn^{2+} > Fe^{2+}$

Reducing agents (list in order of decreasing E^o_{ox}): $Mg > Ni > Sn^{2+} > Fe^{2+} > Cl_2 > Au$

4. a) ox: $[Ag \rightarrow Ag^+ + e^-] \times 2$; red: $Cu^{2+} + 2\,e^- \rightarrow Cu$

balanced: $2\,Ag(s) + Cu^{2+}(aq) \rightarrow 2\,Ag^+(aq) + Cu(s)$

$E^o = E^o_{ox}(Ag) + E^o_{red}(Cu^{2+}) = -0.80\text{ V} + 0.34\text{ V} = -0.46$ V. **Not** spontaneous.

b) ox: $[Ni(s) \rightarrow Ni^{2+} + 2\,e^-] \times 5$; red: $[5\,e^- + 8\,H^+ + MnO_4^- \rightarrow Mn^{2+} + 4\,H_2O] \times 2$

balanced: $5\,Ni(s) + 16\,H^+(aq) + 2\,MnO_4^-(aq) \rightarrow 5\,Ni^{2+}(aq) + 2\,Mn^{2+}(aq) + 8\,H_2O$

$E^o = E^o_{ox}(Ni) + E^o_{red}(MnO_4^-) = +0.23\text{ V} + 1.51\text{ V} = +1.74$ V. Spontaneous.

c) ox: $[2\,H_2O + Mn^{2+} \rightarrow MnO_2 + 4\,H^+ + 2\,e^-] \times 3$

red: $[3\,e^- + 4\,H^+ + NO_3^- \rightarrow NO + 2\,H_2O] \times 2$

balanced: $2\,H_2O + 3\,Mn^{2+}(aq) + 2\,NO_3^-(aq) \rightarrow 3\,MnO_2(s) + 4\,H^+(aq) + 2\,NO(g)$

$E^o = E^o_{ox}(Mn^{2+}) + E^o_{red}(NO_3^-) = -1.21\text{ V} + 0.96\text{ V} = -0.25$ V. **Not** spontaneous.

5. a) $E^o = E^o_{ox}(Fe) + E^o_{red}(Fe^{3+}) = +0.44\text{ V} + 0.77\text{ V} = +1.21$ V. Spontaneous.

b) $E^o = E^o_{ox}(I^-) + E^o_{red}(Zn^{2+}) = -0.54\text{ V} - 0.76\text{ V} = -1.30$ V. **Not** spontaneous.

6. $E^o_{ox}(Br^-) = -1.09$ V. An ion will oxidize Br^- if it has an $E^o_{red} > +1.09$ V.

a) $E^o_{red}(Pb^{2+}) = -0.13\text{ V} \Rightarrow$ No. b) $E^o_{red}(H^+) = 0 \Rightarrow$ No.

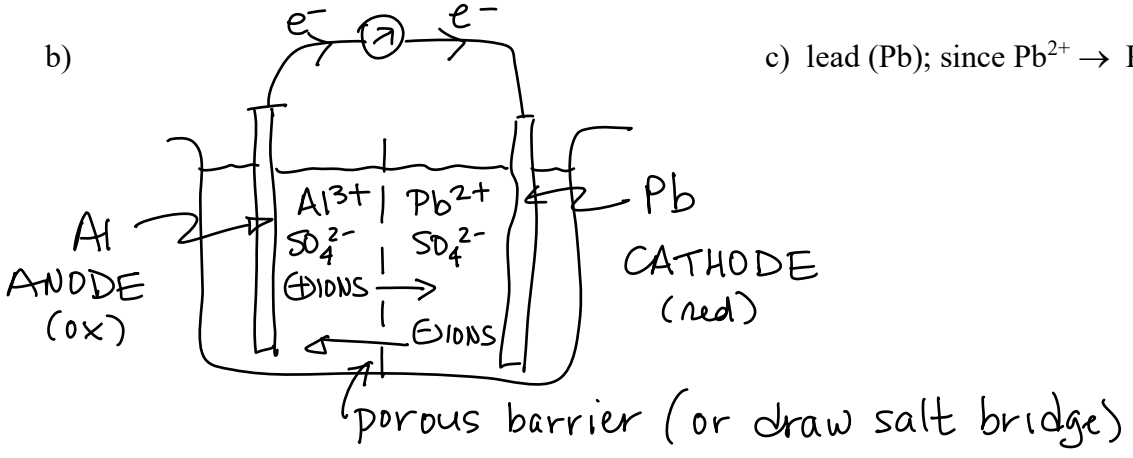
c) $E^o_{red}(Au^{3+}) = +1.50\text{ V} \Rightarrow$ Yes. d) $E^o_{red}(MnO_4^-) = +1.51\text{ V} \Rightarrow$ Yes.

Answer: Au^{3+} and MnO_4^-

7. a) $E^o_{ox}(Al) = +1.66$ V; $E^o_{ox}(Pb) = +0.13$V. Al has a greater tendency to be oxidized. Thus, for spontaneous reaction, Al will be oxidized and Pb will be reduced:

$2\,Al(s) + 3\,Pb^{2+}(aq) \rightarrow 2\,Al^{3+}(aq) + 3\,Pb(s)$

b)



c) lead (Pb); since $Pb^{2+} \rightarrow Pb$

8. a) Iron (Fe) is the strongest reducing agent because it is always oxidized. It reduces all of the other metals.

b) Gold (Au) is the weakest reducing agent because it is never oxidized.

c) Using the voltages of the other metals with Pb.

Half-reaction	E^{o}_{red} (volts)
$Au^{3+} + 3\ e^{-} \leftrightarrow Au$	+ 0.80 v
$Pb^{2+} + 2\ e^{-} \leftrightarrow Pb$	0.00 v
$Ni^{2+} + 2\ e^{-} \leftrightarrow Ni$	– 0.10 v
$Fe^{2+} + 2\ e^{-} \leftrightarrow Fe$	– 0.25 v