Problems of Redox reactions: Galvanic cells

1) A galvanic cell was constructed by connecting the electrodes $Ni^{2+}/Ni_{(s)}$ and $Cr^{3+}/Cr_{(s)}$ where the concentration of each ion is 1 mol/L. Give:

a) Name of each electrode (anode/cathode), sign, half–reactions and type (oxidation/reduction).
b) The net cell reaction.
c) Standard cell voltage at 25 °C.
d) Notation of the galvanic cell if a salt bridge is used.

Data: Standard reduction potentials at 25 °C: $\varepsilon°(Ni^{2+}/Ni_{(s)})=-0.236$ V, $\varepsilon°(Cr^{3+}/Cr_{(s)})=-0.74$ V.

2) A galvanic cell consists of electrodes $Cu^{+}/Cu_{(s)}$ and $NO_3^-/NO_{(g)}$ where the concentration of each ion is 1 mol/L. Gases are in acidic medium at a pressure of 1 atm. Find out:

a) Name of each electrode (anode/cathode), sign, half–reactions and type (oxidation/reduction).
b) The net cell reaction.
c) Standard cell voltage at 25 °C.
d) Notation of the galvanic cell if we use a salt bridge and the gas redox reaction occurs at platinum electrode.

Data: Standard reduction potentials at 25 °C: $\varepsilon°(Cu^{+}/Cu_{(s)})=0.518$ V, $\varepsilon°(NO_3^-/NO_{(g)})=0.96$ V.

3) A galvanic cell consists of electrodes $Al^{3+}/Al_{(s)}$ and $Fe^{2+}/Fe_{(s)}$ where the concentration of each ion is 1 mol/L. Give:

a) Name of each electrode (anode/cathode), sign, half–reactions and type (oxidation/reduction).
b) The net cell reaction.
c) Standard cell voltage at 25 °C.
d) Notation of the galvanic cell if a salt bridge is used.

Data: Standard reduction potentials at 25 °C: $\varepsilon°(Al^{3+}/Al_{(s)})=-1.676$ V, $\varepsilon°(Fe^{2+}/Fe_{(s)})=-0.44$ V.

4) A galvanic cell was constructed by connecting the electrodes $Au^{3+}/Au_{(s)}$ and $Mg^{2+}/Mg_{(s)}$ where the concentration of each ion is 1 mol/L. Give:

a) Name of each electrode (anode/cathode), sign, half–reactions and type (oxidation/reduction).
b) The net cell reaction.
c) Standard cell voltage at 25 °C.
d) Notation of the galvanic cell if a salt bridge is used.

Data: Standard reduction potentials at 25 °C: $\varepsilon°(Au^{3+}/Au_{(s)})=1.42$ V, $\varepsilon°(Mg^{2+}/Mg_{(s)})=-2.357$ V.

Problems of Redox reactions: Galvanic cells

5) A galvanic cell was constructed by connecting the electrodes $Sn^{2+}/Sn_{(s)}$ and $Cr^{3+}/Cr_{(s)}$ where the concentration of each ion is 1 mol/L. Give:

a) Name of each electrode (anode/cathode), sign, half–reactions and type (oxidation/reduction).
b) The net cell reaction.
c) Standard cell voltage at 25 °C.
d) Notation of the galvanic cell if a salt bridge is used.

Data: Standard reduction potentials at 25 °C: $\varepsilon°(Sn^{2+}/Sn_{(s)})=-0.14$ V, $\varepsilon°(Cr^{3+}/Cr_{(s)})=-0.74$ V.

6) A galvanic cell consists of electrodes $Cd^{2+}/Cd_{(s)}$ and $Zn^{2+}/Zn_{(s)}$ where the concentration of each ion is 1 mol/L. Give:

a) Name of each electrode (anode/cathode), sign, half–reactions and type (oxidation/reduction).
b) The net cell reaction.
c) Standard cell voltage at 25 °C.
d) Notation of the galvanic cell if a salt bridge is used.

Data: Standard reduction potentials at 25 °C: $\varepsilon°(Cd^{2+}/Cd_{(s)})=-0.4$ V, $\varepsilon°(Zn^{2+}/Zn_{(s)})=-0.762$ V.

Problems of Redox reactions: Galvanic cells

Answers:

1) **a)** Anode: Oxidation (–): $Cr_{(s)} \rightarrow Cr^{3+} + 3e^-$

Cathode: Reduction (+): $Ni^{2+} + 2e^- \rightarrow Ni_{(s)}$

b) $3Ni^{2+} + 2Cr_{(s)} \rightarrow 3Ni_{(s)} + 2Cr^{3+}$; **c)** 0.504 V;

d) (–) $Cr_{(s)}$ | Cr^{3+} (1 M) || Ni^{2+} (1 M) | $Ni_{(s)}$ (+).

2) **a)** Anode (–): $Cu_{(s)} \rightarrow Cu^+ + e^-$ Oxidation;

Cathode (+): $NO_3^- + 4H^+ + 3e^- \rightarrow NO_{(g)} + 2H_2O$ Reduction;

b) $3Cu_{(s)} + NO_3^- + 4H^+ \rightarrow 3Cu^+ + NO_{(g)} + 2H_2O$; **c)** 0.442 V;

d) (–) $Cu_{(s)}$ | Cu^+ (1 M) || NO_3^- (1 M) | $NO_{(g)}$ (1 atm), Pt (+).

3) **a)** Anode: Oxidation (–): $Al_{(s)} \rightarrow Al^{3+} + 3e^-$

Cathode: Reduction (+): $Fe^{2+} + 2e^- \rightarrow Fe_{(s)}$

b) $3Fe^{2+} + 2Al_{(s)} \rightarrow 3Fe_{(s)} + 2Al^{3+}$; **c)** 1.236 V;

d) (–) $Al_{(s)}$ | Al^{3+} (1 M) || Fe^{2+} (1 M) | $Fe_{(s)}$ (+).

4) **a)** Anode: Oxidation (–): $Mg_{(s)} \rightarrow Mg^{2+} + 2e^-$

Cathode: Reduction (+): $Au^{3+} + 3e^- \rightarrow Au_{(s)}$

b) $2Au^{3+} + 3Mg_{(s)} \rightarrow 2Au_{(s)} + 3Mg^{2+}$; **c)** 3.777 V;

d) (–) $Mg_{(s)}$ | Mg^{2+} (1 M) || Au^{3+} (1 M) | $Au_{(s)}$ (+).

5) **a)** Anode: Oxidation (–): $Cr_{(s)} \rightarrow Cr^{3+} + 3e^-$

Cathode: Reduction (+): $Sn^{2+} + 2e^- \rightarrow Sn_{(s)}$

b) $3Sn^{2+} + 2Cr_{(s)} \rightarrow 3Sn_{(s)} + 2Cr^{3+}$; **c)** 0.6 V;

d) (–) $Cr_{(s)}$ | Cr^{3+} (1 M) || Sn^{2+} (1 M) | $Sn_{(s)}$ (+).

6) **a)** Anode: Oxidation (–): $Zn_{(s)} \rightarrow Zn^{2+} + 2e^-$

Cathode: Reduction (+): $Cd^{2+} + 2e^- \rightarrow Cd_{(s)}$

b) $Cd^{2+} + Zn_{(s)} \rightarrow Cd_{(s)} + Zn^{2+}$; **c)** 0.362 V;

d) (–) $Zn_{(s)}$ | Zn^{2+} (1 M) || Cd^{2+} (1 M) | $Cd_{(s)}$ (+).