Redox Review

1. Define oxidation and reduction.

Leo Ger: lose electron oxidation; gain electron reduction

2. Explain what a redox reaction is.

When electrons are transferred between 2 species; usually from chemical potential to electrical kinetic. Oxidation# +/ -

3. Cu^{2+} reacts spontaneously with Al to produce Cu and Al^{3+}. Write the oxidation half reaction, the reduction half reaction, and the net ionic redox reaction. Label the oxidizing agent and reducing agent.

$Al_{(S)} + Cu^{2+}_{(aq)} <-> Al^{3+}_{(aq)} + Cu_{(s)}$. *O/RA + R/ OA*

4. Using your table, write the half reactions and complete redox reaction when acidified BrO_3^- reacts with H_2S gas. Label the oxidizing agent and reducing agent. Is this reaction spontaneous?

$3H_2S_{(aq)} + HBrO_3^-{}_{(aq)} <-> HBr_{(aq)} + 3S_{(s)} + 3H_2O_{(l)}$ *O/RA + R/ OA*

5. Give the oxidation numbers for each of the following substances:

a) MnO_4^- b) H_2SO_3 c) Fe_3O_4 d) $BaCr_2O_7$ e) C_3H_8 f) $HClO_4$ g) P_4

X=7 x=4 x=8/3 x=6 x=3 x=7 x=0

6. Are the following reactions redox? State why or why not.

a) $BaCl_2 + Na_2SO_4 \Rightarrow BaSO_4 + 2NaCl$

b) $2Na + MgBr_2 \Rightarrow Mg + 2NaBr$

NO change in oxidation #

7. What metal can be oxidized by acidified MnO_4^- but not by acidified BrO_3^-?

Au- see position on table

8. Which is the stronger reducing agent: H_2O_2 or Ni? How do you know?

Ni- see position on table

9. Which substance can be reduced by I^- but not by Fe^{3+}?

MvO_4^- - see position on table

10. If the following reactants are mixed, will the reaction be spontaneous, non spontaneous, or will there be no reaction at all? If spontaneous, write a balanced redox equation:

a) Cu^{2+} and Ag_2S b) K^+ and Sn^{2+} c) $AuCl_4^-$ and Al

NSp-same side NSp-RA above OA Sp- OA above RA

11. Balance the following ~~and calculate the $E°_{cell}$~~:

$Mn^{2+} + ClO_4^- \Rightarrow MnO_4^- + Cl^-$ (acidic)

O/RA. +. R/ OA

O: $(Mn^{2+}_{(aq)} + 4H_2O_{(l)} <-> MnO_{2(s)} + 8H+ + 6e-)\ 4$

R: $(ClO_4^-{}_{(aq)} + 8H^+ + 8e^- <-> Cl^-_{(aq)} + 4H_2O_{(l)})\ 3$

$4Mn^{2+}_{(aq)} + 3ClO_4^-{}_{(aq)} + 4H_2O_{(l)} <-> 4MnO_{2(s)} + 3\ Cl^-_{(aq)} + 8H^+$

12. Balance the following half reaction:

$H_2BO_3^- \Rightarrow BH_4^-$ (basic)

O: $H_2BO_3^-{}_{aq)} + 8H^+ <-> BH_4^-{}_{(aq)} + 3H_2O_{(l)}) + 6e-$

$H_2BO_3^-{}_{aq)} + 5H_2O_{(l)} <-> BH_4^-{}_{(aq)} + 8OH^- + 6e-$

13. Write an oxidation half reaction, reduction half reaction, and overall redox equation for the skeleton redox reaction in basic solution:

$Ag^+ + C_6H_4(OH)_2 \Rightarrow Ag + C_6H_4O_2$

BONUS: Add peroxide!

14. In an unusual compound, IPO_4, iodine exists as iodine (III). The compound decomposes as in the following skeleton redox reaction:

$IPO_4 \Rightarrow I_2 + IO_3^- + PO_4^{3-}$ (acidic)

Balance this redox equation.

O: $(2IPO_4^-{}_{(aq)} + 4H_2O_{(l)} <-> 6H+ + 3e- + 2PO_4^{3-}{}_{(aq)})$ 4

R: $(IPO_4^-{}_{(aq)} + 4e^- <-> IO_{3(aq)}^- + PO_4^{3-}{}_{(aq)})$ 3

$12H_2O_{(l)} + 10IOP_{4(aq)} <-> 3I_{2(s)} + 10\ PO_4^{3-}{}_{(aq)} + 24H^+ + 4IO_3^-{}_{(aq)}$

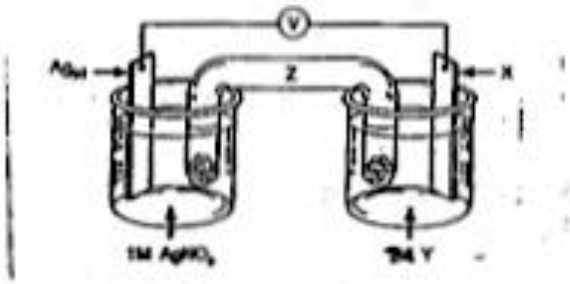
15. In a titration, 28.55mL of acidified 0.0500M $KMnO_4$ is required to oxidize a 10.00mL sample of Cr^{3+}. Write the balanced redox reaction and calculate $[Cr^{3+}]$.

16. In a redox titration, 0.300g of $Na_2C_2O_4$ is placed into a 250mL flask and acidified. The resulting solution requires 23.42mL of $KMnO_4$ to reach the endpoint. The reaction is

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

Using the above data, calculate $[KMnO_4]$.

17. The electrochemical cell below produces an initial voltage of 0.93V.



a) Identify X.
b) Identify a suitable electrolyte Y
c) Identify a suitable electrolyte Z
d) Indicate on the diagram the direction of electron flow.

18. Draw and label a diagram of a cell capable of producing Br_2 from molten NaBr. Label the anode and cathode, then indicate at which electrode Br_2 is produced.