CHM 1046 - JOENS WORKSHEET #9 Due date: Monday, November 20^{th}

WORKSHEETS ARE DUE AT THE BEGINNING OF CLASS ON THE DATE GIVEN ON THE WORKSHEET. LATE WORKSHEETS WILL NOT BE ACCEPTED.

NAME ______________________________ Panther ID ____________________________

For problems involving calculations you must show your work for credit.

1) The solubility product for silver bromide (AgBr, MW = 187.8) is $K_{sp} = 7.7 \times 10^{-13}$ at T = 25. °C.

a) How many grams of AgBr will dissolve in 1.000 L of pure water at T = 25. °C?

The solubility reaction is $AgBr(s) \leftrightarrows Ag^+(aq) + Br^-(aq)$ $K_{sp} = [Ag^+][Br^-] = 7.7 \times 10^{-13}$

	Initial	Change	Equilibrium
Ag^+	0	x	x
Br^-	0	x	x

$(x)(x) = 7.7 \times 10^{-13}$ $x^2 = 7.7 \times 10^{-13}$ $x = (7.7 \times 10^{-13})^{1/2} = 8.8 \times 10^{-7}$ M

So the mass of AgBr that will dissolve in 1.000 L of pure water is

$$\text{mass} = \frac{8.8 \times 10^{-7}\ \text{mol}}{\text{L}}\ \frac{187.8\ \text{g}}{\text{mol}} = 1.65 \times 10^{-4}\ \text{g}$$

b) How many grams of AgBr will dissolve in 1.000 L of a 0.0200 M solution of sodium bromide (NaBr) at T = 25. °C?

We now have a second source of Br^- ion, from NaBr.

Since $NaBr(s) \rightarrow Na^+(aq) + Br^-(aq)$

	Initial	Change	Equilibrium
Ag^+	0	x	x
Br^-	0.0200	x	0.0200 + x

$(x)(0.0200 + x)) = 7.7 \times 10^{-13}$ If we assume $x << 0.0200$, then

$x(0.0200) = 7.7 \times 10^{-13}$ $x = \frac{7.7 \times 10^{-13}}{0.0200} = 3.8 \times 10^{-11}$ M

So the mass of AgBr that will dissolve in 1.000 L of a 0.0200 M solution of NaBr is

$$\text{mass} = \frac{3.8 \times 10^{-11}\ \text{mol}}{\text{L}}\ \frac{187.8\ \text{g}}{\text{mol}} = 7.2 \times 10^{-9}\ \text{g}$$

Note that the presence of a second source of Br^- ion greatly decreases the number of grams of AgBr that will dissolve.

2) Give the oxidation number for sulfur (S) for each of the following substances.

SO_3 ____+6______ H_2S ____-2 ______

$Fe_2(SO_4)_3$ ____+6______ S_8 _____0______

3) Balance each of the following oxidation-reduction reactions for the stated conditions.

a) $S_2O_3^{2-}(aq) + I_2(aq) \rightarrow S_4O_6^{2-}(aq) + I^-(aq)$ (acid conditions)

ox $2\ S_2O_3^{2-}(aq) \rightarrow S_4O_6^{2-}(aq) + 2\ e^-$

red $I_2(aq) + 2\ e^- \rightarrow 2\ I^-(aq)$

net $2\ S_2O_3^{2-}(aq) + I_2(aq) \rightarrow S_4O_6^{2-}(aq) + 2\ I^-(aq)$

This is an unusual reaction because the oxidation number for S in $S_4O_6^{2-}$ is $^5/_2$. Note that in the rare cases where there is a fractional oxidation number, the number of electrons transferred in the corresponding correctly balanced half-reaction will still be an integer value.

b) $Zn(s) + NO_3^-(aq) \rightarrow Zn(OH)_4^{2-}(aq) + NH_3(aq)$ (base conditions)

ox $Zn(s) + 4\ H_2O(\ell) \rightarrow 2\ Zn(OH)_4^{2-} + 4\ H^+(aq) + 2\ e^-$ x 4

red $NO_3^{2-}(aq) + 9\ H^+(aq) + 8\ e^- \rightarrow NH_3(aq) + 3\ H_2O(\ell)$

net (acid) $4\ Zn(s) + NO_3^-(aq) + 13\ H_2O(\ell) \rightarrow 4\ Zn(OH)_4^{2-}(aq) + NH_3(aq) + 7\ H^+(aq)$

convert to base $7\ OH^-(aq) \rightarrow 7\ OH^-(aq)$

net (base) $4\ Zn(s) + NO_3^-(aq) + 7\ OH^-(aq) + 6\ H_2O(\ell) \rightarrow 4\ Zn(OH)_4^{2-}(aq) + NH_3(aq)$

To get the number of electrons to cancel we needed to multiply the oxidation reaction by 4. Also, be sure you understand how the method of adding OH^- ion convert the reaction to balanced for base conditions.