

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  $T = 25.^\circ\text{C}$ .

a) How many grams of AgBr will dissolve in 1.000 L of pure water at  $T = 25.^\circ\text{C}$ ?

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

	Initial	Change	Equilibrium
Ag <sup>+</sup>	0	x	x
Br <sup>-</sup>	0	x	x

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

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  $T = 25.^\circ\text{C}$ ?

We now have a second source of Br<sup>-</sup> ion, from NaBr.

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

	Initial	Change	Equilibrium
Ag <sup>+</sup>	0	x	x
Br <sup>-</sup>	0.0200	x	0.0200 + x

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

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

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<sup>-</sup> 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.

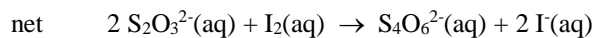
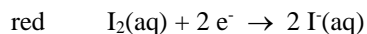
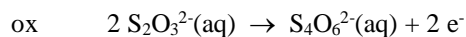
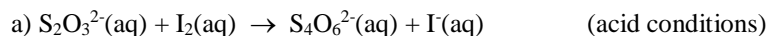
SO<sub>3</sub>      \_\_\_\_\_+6\_\_\_\_\_

H<sub>2</sub>S      \_\_\_\_\_-2\_\_\_\_\_

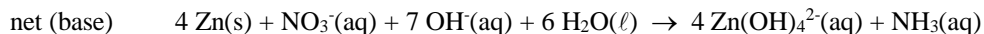
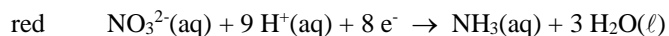
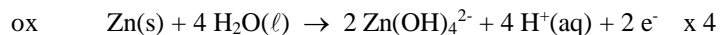
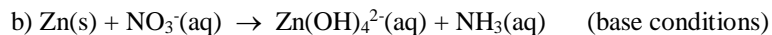
Fe<sub>2</sub>(SO<sub>4</sub>)<sub>3</sub> \_\_\_\_\_+6\_\_\_\_\_

S<sub>8</sub>      \_\_\_\_\_0\_\_\_\_\_

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



This is an unusual reaction because the oxidation number for S in  $\text{S}_4\text{O}_6^{2-}$  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.



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  $\text{OH}^-$  ion convert the reaction to balanced for base conditions.