

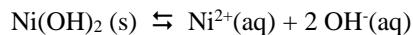
WORKSHEETS ARE DUE AT THE BEGINNING OF CLASS ON THE DATE GIVEN ON THE WORKSHEET. LATE WORKSHEETS WILL NOT BE ACCEPTED.

NAME _____

Panther ID _____

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

1) Using the thermodynamic data given below find the numerical value for the thermodynamic equilibrium constant for the following reaction. You may assume $T = 25.0\text{ }^{\circ}\text{C}$ in doing this problem.



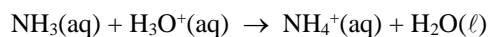
| Substance | ΔH°_f (kJ/mol) | ΔG°_f (kJ/mol) | S° (J/mol·K) |
|------------------------------------|-------------------------------|-------------------------------|-----------------------|
| $\text{Ni}^{2+}(\text{aq})$ | - 64.0 | - 46.4 | 30.1 |
| $\text{Ni}(\text{OH})_2(\text{s})$ | - 538.06 | - 453.1 | 79.5 |
| $\text{OH}^{-}(\text{aq})$ | - 229.94 | - 157.30 | - 10.5 |

$$\begin{aligned} \Delta G^{\circ}_{\text{rxn}} &= [\Delta G^{\circ}_f(\text{Ni}^{2+}(\text{aq})) + 2 \Delta G^{\circ}_f(\text{OH}^{-}(\text{aq}))] - [\Delta G^{\circ}_f(\text{Ni}(\text{OH})_2(\text{s}))] \\ &= [(- 46.4) + 2 (- 157.30)] - [(- 453.1)] = + 92.1 \text{ kJ/mol} \end{aligned}$$

$$\ln K = - \frac{\Delta G^{\circ}_{\text{rxn}}}{RT} = - \frac{(92100. \text{ J/mol})}{(8.314 \text{ J/mol}\cdot\text{K})(298. \text{ K})} = - 37.17$$

$$K = e^{-37.17} = 7.2 \times 10^{-17}$$

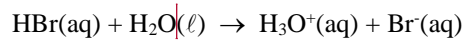
2) For the following reaction identify the Bronsted acid, the Bronsted base, the conjugate base of the Bronsted acid, and the conjugate acid of the Bronsted base.

Bronsted acid _____ H_3O^{+} _____Bronsted base _____ NH_3 _____Conj base of
Bronsted acid _____ H_2O _____Conj acid of
Bronsted base _____ NH_4^{+} _____

3) Find the value for pH for the following solutions at T = 25.0 °C

a) A 0.044 M solution of hydrobromic acid (HBr), a strong acid

A strong acid, so

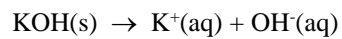


$$[\text{H}_3\text{O}^+] = \frac{0.044 \text{ mol HBr}}{1 \text{ L}} \frac{1 \text{ mol H}_3\text{O}^+}{1 \text{ mol HBr}} = 0.044 \text{ M H}_3\text{O}^+$$

$$\text{pH} = -\log_{10} [\text{H}_3\text{O}^+] = -\log_{10} (0.044) = 1.36$$

b) A 0.128 M solution of potassium hydroxide (KOH), a strong soluble base

A strong soluble base, so



$$[\text{OH}^-] = \frac{0.128 \text{ mol KOH}}{1 \text{ L}} \frac{1 \text{ mol OH}^-}{1 \text{ mol KOH}} = 0.128 \text{ M OH}^-$$

$$\text{pOH} = -\log_{10} [\text{OH}^-] = -\log_{10} (0.128) = 0.89$$

$$\text{pH} = 14.00 - \text{pOH} = 14.00 - 0.89 = 13.11$$