<u>Chemistry 534</u>

<u>A-</u> <u>Extra Arrhenius and Bronsted-Lowry Practice</u>
<u>B-</u> <u>Intro to pH and pOH</u>

1. For each equilibrium, identify one Bronsted-Lowry base for the forward reaction, and do likewise for the reverse reaction.

In World War II **Johannes Nicolaus Brønsted** opposed the Nazis, and in consequence he was elected to the Danish parliament in 1947, but could not take his seat because of illness. Shortly after the election, he died.

a) $CH_3COOH(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CH_3COO^-(aq)$

b) $H_2O(l) + NH_3(aq) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$

2. Given: $HCO_3^-(aq) + H_2O(l) \rightleftharpoons H_2CO_3(aq) + OH^-(aq)$

a) Identify two Bronsted Lowry acids in the following equilibrium.
b) CO_2, when added to water, forms H_2CO_3. If the thickness of a chicken egg shell is proportional to the amount of HCO_3^- in its blood stream, explain whether or not it makes sense to give carbonated water to chickens. (Assuming of course, that you do want thicker egg shells.)

3. Show how $H_2PO_4^{-2}$ could act as a Bronsted-Lowry base in water.

4. Write an ionic equation to show how $Mg(OH)_2$ acts as an Arrhenius base.

5. a) How many moles of H^+ will result if 3 moles of H_2SO_4 completely dissociate?
b) H^+ does not really exist as a separate species in water? What then is the true acidic ion produced by something like HCl in water?

6. ***Recall*** **:**

pH = -log[H$^+$],	where [H$^+$] = moles/L of H$^+$
pOH = -log[OH$^-$],	where [OH$^-$] = moles/L of OH$^-$
pH + pOH = 14	at 25^{0}C

a) If 0.034 g of OH^- are dissolved in 10.0 L of water, what is the pH of the solution?
b) Which has a lower pH: A solution with 10^{-8} moles/L of H^+? Or a solution with 10^{-9} moles/L of OH^-? Show why.
c) If the pOH keeps decreasing, what is happening to the amount of OH^- dissolved in that solution?
d) If pH keeps increasing, what is happening to the amount of H^+ or H_3O^+ dissolved in that solution?

7. **$pH = -log[H^+]$** can also be written as $[H^+] = 10^{-pH}$.

a) If clean rain is supposed to have a pH of about 5.6, how many times more acidic(how many times more $[H^+]$) is a rain sample with a pH of 4.2, close to the average pH of rain falling over Montreal?

b) How many times stronger is an NaOH with pOH = 2.50 solution compared to one at 2.75?

8. Use the laws of logarithms and the formula from last year : $[H^+]\ [OH-] = 10^{-14}$ to derive **$pH + pOH = 14$**

Answers

1. a) **Bronsted bases**: H_2O, CH_3COO^-

   b) **Bronsted bases**: NH_3, OH^-

2. $$HCO_3^-(aq) + H_2O(l) \rightleftharpoons H_2CO_3(aq) + OH^-(aq)$$

   a) H_2O and H_2CO_3
   b) Yes, adding CO_2 ,which will increase H_2CO_3, makes sense. It will encourage the reverse reaction and increase the concentration of HCO_3^-.

3. $H_2PO_4^{-2} + H_2O \rightleftharpoons H_3PO_4 + OH^-$

4. $Mg(OH)_2 \rightleftharpoons Mg^{+2} + 2\ OH^-$

5. a) $H_2\ SO_4 \rightleftharpoons 2\ H^+ + SO_4^{-2}$ (assuming complete dissociation, which will only happen if a base forces it to do so)
   3 moles of $H_2SO_4(2\ H^+/1\ H_2SO_4) = 6$ moles H^+
   b) H_3O^+ is the true acidic species (it's your dad, not Santa Claus)

6. a) **pOH = -log(0.034g(mole/17g)/10 L) = 3.70; pH = 14 - 3.70 = 10.3**
   b) 10^{-8} moles/L of H^+: pH = 8; other one has a pOH of 9 but a pH of 5, and so is more acidic
   c) The concentration of hydroxide keeps increasing with decreasing pOH.
   d) If pH keeps increasing, the amount of H^+ or H_3O^+ keeps decreasing.

7. a) $10^{-4.2}/ 10^{-5.6} = 25.1$ times stronger

   b) $10^{-2.5}/ 10^{-2.75} = 1.78$ times stronger

8. $[H^+]\ [OH-] = 10^{-14}$

   Log both sides:

   $\log[H^+] + \log\ [OH-] = \log 10^{-14}$

   Multiply equation by -1:

   $-\log[H^+] - \log\ [OH^-] = -\log 10^{-14}$

   pH + pOH = 14