This is ACA # 20. It is OK to use your textbook, but if you can answers the questions without it that is OK too.

I recommend you print out this page and bring it to class. Click here to show a set of five ACA20 student responses, randomly selected from all of the student responses thus far, in a new window.

John , here are your responses to the ACA and the Expert's response.

1. The chemical equation that describes how HCl(aq) behaves as a Bronsted-Lowry acid is

$$HCl(aq) + H_2O(l) \rightleftarrows H_3O^+(aq) + Cl^-(aq)$$

In this reaction HCl donates a proton to H_2O, as evidenced by the presence of $Cl^-(aq)$ as a product (HCl must donate a proton to become Cl^-) and $H_3O^+(aq)$ (H_2O must accept a proton to become H_3O^+). So identifying everyone in the reaction as an acid or base in Bronsted-Lowry terms,

$HCl(aq)$	$+ H_2O(l)$	$\rightleftarrows$	$H_3O^+(aq)$	$+ Cl^-(aq)$
acid	**base**		**conjugate acid**	**conjugate base**

For each of the following acids write the Bronsted-Lowry equation that describes its acidic character. (NOTE: It is a little difficult to represent the equilibrium arrow online so using a ---> (three '-' dashes and a '>' greater than character is fine. Just remember for weak acids the equilibrium symbol is best for representing the reaction arrow.)

a) $HClO_4(aq)$

no charge 9%
Arrhenius: 12%
wrong formula: 9%

HClO4(aq) + H2O(l) --> H3O^+(aq) + ClO4^-(aq) 67%

$HClO_4(aq) + H_2O(l) \text{--->} H_3O^+(aq) + ClO_4^-(aq)$

$HClO_4(aq) + H_2O(l) \rightleftarrows H_3O^+(aq) + ClO_4^-(aq)$

acid base conjugate acid conjugate base

b) $HC_2H_3O_2(aq)$

HC2H3O2(aq) + H2O(l) --> H3O^+(aq) + C2H3O2^-(aq)

$HC_2H_3O_2(aq) + H_2O(l) \longrightarrow H_3O^+(aq) + C_2H_3O_2^-(aq)$

$HC_2H_3O_2(aq) + H_2O(l) \rightleftarrows H_3O^+(aq) + C_2H_3O_2^-(aq)$

acid base conjugate acid conjugate base

c) $CH_3NH_3^+(aq)$

CH3NH3^+(aq) + H2O(l) --> H3O^+(aq) + CH3NH2(aq)

$CH_3NH_3^+(aq) + H_2O(l) \longrightarrow H_3O^+(aq) + CH_3NH_2(aq)$

$HC_2H_3O_2(aq) + H_2O(l) \rightleftarrows H_3O^+(aq) + C_2H_3O_2^-(aq)$

acid base conjugate acid conjugate base

2. The chemical equation that describes how $NH_3(aq)$ behaves as a Bronsted-Lowry base is

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

In this reaction NH_3 accepts a proton to H_2O, as evidenced by the presence of $NH_4^+(aq)$ as a product (NH_3 must accept a proton to become NH_4^+) and $OH^-(aq)$ (H_2O must donate a proton to become OH^-). So identifying everyone in the reaction as an acid or base in Bronsted-Lowry terms,

$NH_3(aq)$	$+ H_2O(l)$	$\rightleftarrows$	$NH_4^+(aq)$	$+ OH^-(aq)$
base	acid		conjugate acid	conjugate base

For each of the following bases write the Bronsted-Lowry equation that describes its basic character.

a) $Cl^-(aq)$

no charge: 3%
?: 28%
acid: 3%

Cl^-(aq) + H2O(l) --> HCl(aq) + OH^-(aq) 62%

$Cl^-(aq) + H_2O(l) \text{ ---> } HCl(aq) + OH^-(aq)$

$Cl^-(aq)$	$+ H_2O(l)$	$\rightleftarrows$	$HCl(aq)$	$+ OH^-(aq)$
base	acid		conjugate acid	conjugate base

b) $HCO_3^-(aq)$

HCO3^-(aq) + H2O(l) --> H2CO3(aq) + OH^-(aq)

$HCO_3^-(aq) + H_2O(l) \text{ ---> } H_2CO_3(aq) + OH^-(aq)$

$HCO_3^-(aq)$	$+ H_2O(l)$	$\rightleftarrows$	$H_2CO_3(aq)$	$+ OH^-(aq)$
base	acid		conjugate acid	conjugate base

3. Complete the following table by calculating the value of the missing entry. (NOTE: To express a number in scientific notation 1.00×10^{-7}, enter the value as 1.00e-7.)

$[H^+]$	$[OH^-]$	pH	pOH

5.50×10^{-5} M	1.8e-10 M 1.82×10^{-10} M 90%	4.26 4.26 90%	9.74 9.74 90%
1.4e-6 M 1.42×10^{-6} M 85%	7.06×10^{-9} M	5.85 5.85 85%	8.15 8.15 85%
1.26e-12 M 1.26×10^{-12} M	7.94e-3 M 7.94×10^{-3} M	11.9 11.90	2.10
4.68e-7 M 4.68×10^{-7} M	2.14e-8 M 2.14×10^{-8} M	6.33	7.67 7.67

In the first row the $[H^+] = 5.50 \times 10^{-5}$ M

to calculate the $[OH^-]$ we use the equilibrium expression for water

$$K_w = 1.00 \times 10^{-14} = [H^+][OH^-]$$

$1.00 \times 10^{-14} = [H^+][OH^-]$

substitute $[H^+] = 5.50 \times 10^{-5}$ M and solve for $[OH^-]$

$1.00 \times 10^{-14} = 5.50 \times 10^{-5}$ M$[OH^-]$

$[OH^-] = 5.50 \times 10^{-5}$ M$/1.00 \times 10^{-14} = 1.82 \times 10^{-10}$ M

To calculate the pH of the solution

$$pH = - \log [H^+]$$

$pH = - \log (5.50 \times 10^{-5}) = 4.26$

To calculate the pOH of the solution

$$14 = pH + pOH \text{ or } pOH = -\log [OH^-]$$

$14 = pH + pOH = 2.60 + pOH$

$pOH = 14 - 4.23 = 9.74$

or

$pOH = - \log (1.82 \times 10^{-10}) = 9.74$

The calculation in the second row are similar.

In the third row we need to be able to do an antilog. We are given the pOH and we must calculate the $[H^+]$. Since

$pOH = -\log[OH^-]$

$2.10 = -\log[OH^-]$

We must first move the negative sign to the other side of the equal sign,

$-11.10 = \log[OH^-]$

raise each side to the power of 10

$10^{-2.10} = 10^{-\log[OH^-]}$

In our calculator use the 10^x function to calculate $10^{-2.10}$

$[OH^-] = 10^{-11.10} = 7.94 \times 10^{-3}$ M

We can calculate the $[H^+]$ as above.

The fourth row is calculated similar to the third row.

4. Calculate the pH of 0.250 M $HC_2H_3O_2$ solution. strong acid: 29%

pH = 2.67 38% ~~X~~ 33%

To calculate the pH of the solution we have to know the $[H^+]$ at equilibrium. Assuming $HC_2H_3O_2$ behaves as a weak acid we can set up the ICE table,

	$HC_2H_3O_2(aq)$	$\rightleftarrows H^+(aq)$	$+ C_2H_3O_2^-(aq)$
I	0.250 M	~0	0
C	-x	+x	+x
E	0.250 - x	+x	+x

The equilibrium expression for the reaction is $Ka = [H^+][C_2H_3O_2^-]/[HC_2H_3O_2]$ the value of Ka is 1.8×10^{-5}

Substituting,

$1.8 \times 10^{-5} = [H^+][C_2H_3O_2^-]/[HC_2H_3O_2]$

$1.8 \times 10^{-5} = (x)(x)/(0.250 - x)$

assume that x <<<<0.250

$1.8 \times 10^{-5} = (x)(x)/(0.250)$

$4.5 \times 10^{-6} = x^2$

2.1×10^{-3} M = x = [H^+]

so the pH is 2.67

5. Calculate the pH of 0.250 M NH_3 solution.

strong base 10%

x 52%

pH = 11.33

38%

To calculate the pH of the solution we have to know the [OH^-] at equilibrium. Assuming NH_3 is a weak base we can set up the ICE table,

	NH_3(aq)	+ H_2O	$\rightleftarrows$ OH^-(aq)	+ NH_4^+(aq)
I	0.250 M	-	~0	0
C	-x	-	+x	+x
E	0.250 - x	-	+x	+x

The equilibrium expression for the reaction is Kb = [OH^-][NH_4^+]/[NH_3] the value of Kb is 1.8×10^{-5}

Substituting,

1.8×10^{-5} = [OH^-][NH_4^+]/[NH_3]

1.8×10^{-5} = (x)(x)/(0.250 - x)

assume that x <<<<0.250

1.8×10^{-5} = (x)(x)/(0.250)

$4.5 \times 10^{-6} = x^2$

2.1×10^{-3} M = x = [OH^-]

so the pOH is 2.67 and pH = 11.32

6. Is there anything about the questions that you feel you do not understand? List your concerns/questions.

nothing

7. If there is one question you would like to have answered in lecture, what would that question be?

nothing