PART I

1. Write the equilibrium expression for the following:
   $CH_3COOH(aq) + H_2O(l) \longleftrightarrow CH_3COO^- (aq) + H_3O^+ (aq)$
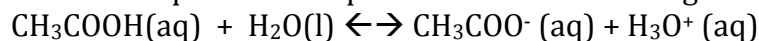
$$K = \frac{[CH_3COO^-][H_3O^+]}{[CH_3COOH]}$$

2. Given that K = 1.8 X 10^{-5} for this reaction, does it favor products or reactants?
   K<1, therefore reactants are favored.

3. Decide whether $\Delta G°_{reaction}$ is "+" or "-".
   If K<1, $\Delta G°_{reaction}$>0, "+," that is the reaction is nonspontaneous
   If you are unsure, you can confirm this with the all-important equation relating
   ΔG reaction to K:

   $$\Delta G°_{reaction} = -RTln(K)$$
   At standard conditions R = 8.314 J/molK and T = 298 K
   $\Delta G°_{reaction}$ = -(8.314 J/molK)(298K)(ln(1.8 X 10^{-5}))
   $\Delta G°_{reaction}$ = 27068 J/mol or about 27.1 KJ/mol

   This is a large, positive $\Delta G°_{reaction}$ as expected by such a small K.

4. If you start with 0.2 M concentration of CH_3COOH and no products, do you
   expect to have a higher concentration of CH_3COOH or CH_3COO^- at
   equilibrium?
   There will still be more reactants, a higher concentration of CH_3COOH at the end, although
   there will be some products as well.

5. Start with a 0.2 molar solution of CH_3COOH and assume 0 concentration of
   the products. What will be the concentrations of all the species in the
   equilibrium mixture given that the K = 1.8 x 10^{-5} for this reaction as written.
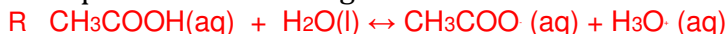
R	$CH_3COOH(aq)$ + $H_2O(l)$ ↔	CH_3COO^- (aq) +	H_3O^- (aq)
I	0.2M	0	0
C	-x	+x	+x
E	0.2M - x	x	x

$$K = \frac{[CH_3COO^-][H_3O^+]}{[CH_3COOH]}$$

$$K = \frac{x^2}{0.2 - x}$$

Since our value for K is so small (smaller than 10^{-3}), we can assume x is negligible

$$1.8 \times 10^{-5} = \frac{x^2}{0.2}$$

$$(1.8 \times 10^{-5}) \times 0.2 = x^2 \quad 1.9$$

$$x = \sqrt{3.6 \times 10^{-6}}$$

$$x = 1.9 \times 10^{-3}$$

[CH₃COOH] = 0.2-1.9x10⁻³ = 0.198 M
[CH₃COO] = 1.9x10⁻³ M
[H₃O⁻] = 1.9x10⁻³ M

**CLICKER QUESTION – BE PREPARED TO ENTER YOUR ANSWERS FOR #5.**

PART II

READ THIS IMPORTANT INFORMATION:

Strong acids and weak acids both dissociate or ionize in water to form protonated water, hydronium ion, and the anion of the acid (what is left of the molecule after the "acidic" proton is removed from the rest of the molecule.) Strong bases dissociate to form a cation and the hydroxide ion (capable of accepting a proton) in solution, whereas weak bases are typically nitrogen-containing molecules that will accept a proton. Before we can begin to study acid/base chemistry we have to be able recognize compounds as acids or bases as strong or weak and understand how they behave in aqueous solutions.

The difference is where the equilibrium lies … far to the left or far to the right.

Strong acids lie so far to the right that one can assume, for dilute solutions of strong acids, the extent of ionization is 100%. For example, the K for HCl(aq) = 1.3 X 10⁶
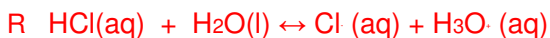
1. Calculate the concentration of hydronium ion for the dissociation reaction for the strong acid hydrochloric acid, given a starting concentration of 0.2 molar HCl. Assume 100% ionization.

$$HCl(aq) + H_2O(l) \rightarrow Cl^-(aq) + H_3O^+(aq)$$
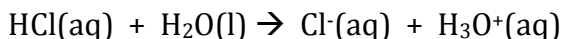
If the reaction goes 100%, we should have 100% products in the end with 0 reactants

Since there is no change in volume, and everything in the reaction is in a 1:1 stochiometric ratio, we can leave our units in terms of M

IF you have a reaction where the volumes are changing, it is *BEST* to work your RICE table in moles.

R   HCl(aq) + H₂O(l) ↔ Cl⁻ (aq) + H₃O⁺ (aq)
I   0.2M   0   0
C   -0.2M   +0.2M   +0.2M
E   0M   0.2M   0.2M

[H₃O⁻] = 0.2M

2. Compare the concentration of hydronium ion from the dissociation of acetic acid, to the dissociation of hydrochloric acid. Although, the acids started with

the same molar concentration of un-dissociated compound, the amount of hydronium ion is quite different. Please explain this observation.

The [H3O⁺] in the HCl solution is much higher than in the acetic acid solution (0.2M > 1.9x10⁻³M)

HCl is a strong acid and so it completely dissociates in water, it has a large K value. Essentially all of its reactant acidic protons leave the HCl compound to form H3O⁺ product. Acetic acid has a small K value so only a few of the reactant acidic protons on acetic acid leave the compound to form the product H3O⁺ in solution. That is, there are more reactants than products at equilibrium. The concentration of H3O⁺ in solution is therefore much higher in the HCl solution in which essentially 100% of the acidic protons are forming H3O⁺ product.

PART III
Demonstration
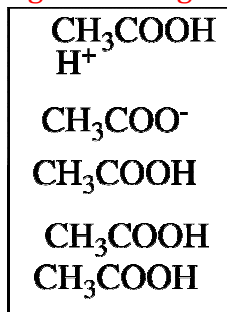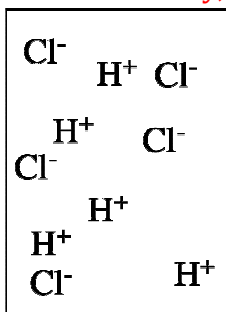Putting Acetic Acid and Hydrochloric Acid in Water.

Macro and Micro description here please:
With pure water, the light bulb is not lit
With 1M acetic acid, the light bulb IS lit
With 1M HCl, the light bulb is VERY bright
It is the concentration of ions that causes the electricity to flow. The more ions, the more electricity, the brighter the light.

Include Indicator

The reactant the chemical equation in PART I question (1) is called Acetic Acid. Acetic acid is considered a weak acid because it's equilibrium lies far to the left. This means that when placed in water, only a small % of the compound dissociates or ionizes to form the anion of the acid and protonated water, or hydronium ion.

Strong Electrolyte vs Weak Electrolyte

Acid Base Brain Dump

1. Think about everything you know about acids and bases and write it down here. Be prepared to share a statement with the class if called upon.
Acids: Corrosive, sour, citrus, stomach acid, HCl, H_2SO_4, vinegar, Red on Litmus paper

The Bronsted-Lowry definition of an acid is that the compound is a proton donor, and the Bronsted-Lowry definition of a base is that the compound is a proton acceptor. Each time a proton is donated across a chemical change, one compound is acting like an acid and another is acting like a base.
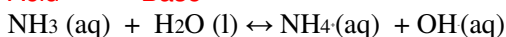
2. Label what is acting like an acid and what is acting like a base in the following reactions:
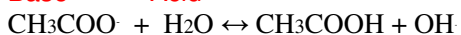
$CH_3COOH(aq) + H_2O(l) \leftrightarrow CH_3COO^-(aq) + H_3O^+(aq)$
Acid   Base

$HCl(aq) + H_2O(l) \rightarrow Cl^-(aq) + H_3O^+(aq)$
Acid   Base

$NH_3(aq) + H_2O(l) \leftrightarrow NH_4^+(aq) + OH^-(aq)$
Base   Acid

$CH_3COO^- + H_2O \leftrightarrow CH_3COOH + OH^-$
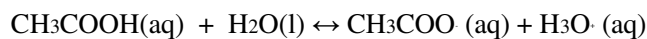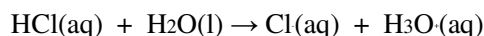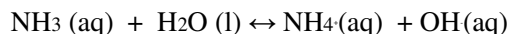Base   Acid

WHOLE CLASS CHECK IN… WHOLE CLASS CHECK IN

In the context Bronsted-Lowry acids and bases, if an acid transfers a proton the "deprotonated" compound is referred to as the "conjugated base". Conversely an base that accepts a protons becomes a "conjugate acid" upon protonation. For the reactions below identify the conjugate acids and bases in the products.
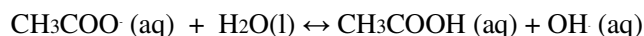
$CH_3COOH(aq) + H_2O(l) \leftrightarrow CH_3COO^-(aq) + H_3O^+(aq)$
                           C. BaseC. Acid

$HCl(aq) + H_2O(l) \rightarrow Cl^-(aq) + H_3O^+(aq)$
                   C. BaseC. Acid

$NH_3(aq) + H_2O(l) \leftrightarrow NH_4^+(aq) + OH^-(aq)$
                   C. Acid C. Base

$CH_3COO^-(aq) + H_2O(l) \leftrightarrow CH_3COOH(aq) + OH^-(aq)$
                  C. Acid   C. Base