# Acid and base worksheet 4      Name _______________

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

$pH + pOH = 14$

Acids have a dissociation

$$HA = H^+ + A^-$$

**Weak Acids** do not completely dissociate. If a solution of a weak acid is 0.10 M HA this does not mean that the $[H^+]$ will be 0.10 M. You must consider the $K_a$ and determine how much $H^+$ dissociated.

$$K_a = \frac{[H^+][A^-]}{[HA]}$$

To determine the pH we need to know the $[H^+]$.

If the $K_a = 7.1x10^{-4}$ of a 0.10M $HNO_2$, what is the pH of the acid?

| | HA | = | $H^+$ | + | $A^-$ |
|---|---|---|---|---|---|
| Initial concentration | 0.10 | | 0 | | 0 |
| Change | -x | | +x | | +x |
| Equilibrium | 0.10-x | | +x | | +x |

$A^-$ represents $NO_2^-$

Setup the equilibrium expression. $$7.1x10^{-4} = \frac{[x][x]}{[0.10 - x]}$$

Solve for x      $7.1x10^{-4}(0.10\text{-}x) - x^2 = 0$

Graph and find x      x = 0.0081

Looking back at the table above x is equal to the concentration of $H^+$ and $A^-$ at equilibrium.

Therefore the pH will equal the      **pH = -log[0.0081] = 2.1**

Calculate the following

1. What is the pH of a 0.25 M solution of HF if the $K_a = 6.8x10^{-4}$?

_____________________ans.