

Acid and base worksheet 4

Name _____

$$\text{pH} = -\log [\text{H}^+]$$

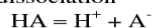
$$\text{pOH} = -\log [\text{OH}^-]$$

$$[\text{H}^+] = 10^{-\text{pH}}$$

$$[\text{OH}^-] = 10^{-\text{pOH}}$$

$$\text{pH} + \text{pOH} = 14$$

Acids have a dissociation



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

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

To determine the pH we need to know the $[\text{H}^+]$.

If the $K_a = 7.1 \times 10^{-4}$ of a 0.10M HNO_2 , what is the pH of the acid?

	HA	=	H ⁺	+	A ⁻	
Initial concentration	0.10		0		0	A ⁻ represents NO ₂ ⁻
Change	-x		+x		+x	
Equilibrium	0.10-x		+x		+x	

Setup the equilibrium expression. $7.1 \times 10^{-4} = \frac{[x][x]}{[0.10 - x]}$

Solve for x $7.1 \times 10^{-4}(0.10-x) - x^2 = 0$

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

- What is the pH of a 0.25 M solution of HF if the $K_a = 6.8 \times 10^{-4}$?

_____ans.