How to Determine the pH of Weak Acids and Bases.

1. Read the problem.
2. Determine if the acid is strong or weak.
3. For strong acids, the hydronium ion concentration is the same as the acid concentration, because there is 100% dissociation of the proton from the acid.
4. For weak acids, there is an equilibrium between the weak acid and its conjugate base. So the equilibrium expression must be used to determine hydronium ion concentration.
5. Use the ICE method to determine the equilibrium concentration of hydronium ions.

a. Look up the Ka
b. Write the balanced equation.
c. Filll in the ICE table
d. Solve for $[H_3O^+]$
e. pH = $-\log[H_3O^+]$

1. Sample problem: What is the pH of a 0.02M solution of carbonic acid?

2. Is this a weak acid or a strong acid?

3. For strong: pH = $-\log[H_3O^+]$

4. For weak: Ka = $\dfrac{[H_3O^+][A^-]}{[HA]}$ let $[H_3O^+]$ = x, x also = A- (see balanced equation)

5 a. Ka for carbonic acid is 1.7×10^{-4}

b. Balanced Eq:	H_2CO_3	çè	HCO_3^-	H_3O^+
c. Initial (get from problem)	0.02		0	0
Change (x number of moles of HA lose their protons to become A-)	-x		x	x
Equilibrium (sum of I and C)	0.02-x		x	x

Plug the equilibrium concentrations into the Ka expression and solve for x. Since x = $[H_3O^+]$, take the negative log of x to get the pH.

d. $1.7 \times 10^{-4} = \dfrac{x * x}{0.02-x}$

If x is less than 5% of the concentration of the weak acid, it can be ignored in the denominator, as it is subtracted. A large number minus a very small number is essentially equal to the large number. For example, if you have 100,000 pennies and 4 pennies get lost, you still have about a $1,000.

Ignor x in the denominator: $1.7 \times 10^{-4} = \frac{x * x}{0.02}$

$0.02 * 1.7 \times 10^{-4} = x^2$

x = squareroot of 3.4×10^{-6}

$x = 0.0018$

e. pH = -log x   pH = -log .0018   pH = 2.74

Now you try it: What is the pH of a 0.05M solution of sodium dihydrogen phosphate? ($K_a = 1.38 \times 10^{-7}$)