

Acid Base Worksheet

Arrhenius's defines acids as substances when placed in water produced hydronium (H_3O^+). He defines bases as substances when placed in water produce hydroxide ions (OH^-).

Strong acids and bases are substances that have nearly complete dissociation. If a 0.10 Molar solution of HCl is created the hydronium concentration will also be 0.10 M due to complete dissociation. The pH is defined as the negative logarithm of the hydronium concentration. The pH of the hydrochloric solution would be calculated by the following equation (eq. 1.).

$$\text{eq. 1 } \text{pH} = -\log([\text{H}^{(+)}])$$

The pOH on the other hand is the negative logarithm of the hydroxide concentration.

$$\text{eq. 2 } \text{pOH} = -\log([\text{OH}^{(-)}])$$

The sum of the pH and the pOH is 14. If there is a pH of 1 then it would have a pOH of 13.

$$\text{eq. 3 } \text{pH} + \text{pOH} = 14$$

Weak acids and bases do not completely dissociate. The dissociation constant for an acid is a ratio of the hydronium times the anion divided by the concentration of the acid at equilibrium as demonstrated by the following equation.

$$\text{eq. 4 } K_a = \frac{[\text{H}^{(+)}][\text{A}^{(-)}]}{[\text{HA}]}$$

What is the pH of a 0.10 Molar solution of acetic acid? The pH will still be the negative log of the hydronium (H^+) concentration but the hydronium concentration in a weak acid is not equal to the initial concentration of the acid. Acetic acid has a $K_a = 1.8 \times 10^{-5}$. To calculate the H^+ concentration the change in the initial concentration must be considered. In the table below the initial concentration of the acid is 0.10 M. When it is placed in water some of it breaks up into H^+ and A^- and some remains as HA at equilibrium.

	HA	->	H+	A-
Initial Concentration	.10		0	0
Change in Concentration	-x		+x	+x
Final Concentration	.10-x		x	x

The calculation then is $1.8\text{E-}5 = \frac{(x) \cdot (x)}{.10 - x}$. x is the concentration of H^+ at equilibrium.

$$\text{solve} \left(1.8\text{E-}5 = \frac{(x) \cdot (x)}{(0.10 - x)}, x \right) \quad x = 1.3\text{e-}3 \text{ or } x = -1.4\text{e-}3$$

There are two solutions for this equation, but only one right answer to the problem. There can not be a negative concentration for H⁺, therefore 1.3×10^{-3} is the proper Molar concentration of H⁺ and A⁻. From this the pH can be determined by using *eq 1*.

1. Calculate the following data for **strong** Acids and Bases.

	Molarity	[H ⁺]	pH	[OH ⁻]	pOH
HCl			2.00		
HNO ₃		.00250			
HI	.0500				
NaOH	.0500				
KOH			9.25		
LiOH				.00450	

2. Calculate the following data for **weak acids** and bases

	ka	Molarity	[H ⁺]	pH
HC ₂ H ₃ O ₂	1.80×10^{-5}	.100		
HCNO	3.50×10^{-4}			3.00
HF		.100		

3. Determine the conjugate of each compound.

	Conjugate
HCN	
HC ₂ H ₃ O ₂	
NH ₃	
SO ₄ ⁻²	
N ₂ H ₄	

4. Write the dissociation equation for the following acids.

- A. $\text{HCl} + \text{H}_2\text{O} \leftrightarrow$
 B. $\text{NH}_4^{(+)} + \text{H}_2\text{O} \leftrightarrow$
 C. $\text{HSO}_4^{(-)} + \text{H}_2\text{O} \leftrightarrow$
 D. $\text{HC}_2\text{H}_3\text{O}_2 + \text{H}_2\text{O} \leftrightarrow$

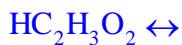
5. Write the dissociation equation for the following bases.

- A. $\text{C}_6\text{H}_5\text{NH}_2 + \text{H}_2\text{O} \leftrightarrow$
 B. $\text{NH}_2\text{OH} + \text{H}_2\text{O} \leftrightarrow$
 C. $(\text{CH}_3)_3\text{N} + \text{H}_2\text{O} \leftrightarrow$

6. A buffered solution is created from a weak acid with its conjugate salt or a weak base with its conjugate salt.

Determine the pH of a solution made of 0.100 M acetic acid ($\text{HC}_2\text{H}_3\text{O}_2$) and 0.100 M sodium acetate ($\text{NaC}_2\text{H}_3\text{O}_2$). The K_a for acetic acid is 1.80×10^{-5} .

Write the dissociation equation for both the acetic acid and the sodium acetate



	HC ₂ H ₃ O ₂	H ⁺	C ₂ H ₃ O ₂ ⁻
Initial	.100	.000	.100
Change	-x	+x	+x
Final	0.100-x	+x	0.100+x

(Note: the 0.100 M in the initial $\text{C}_2\text{H}_3\text{O}_2^-$ box is from the concentration of the sodium acetate added)

$$K_a = \frac{[\text{H}^{(+)}][\text{C}_2\text{H}_3\text{O}_2^{(-)}]}{[\text{HC}_2\text{H}_3\text{O}_2]}$$

1. Calculate the approximate pH of a solution that contains 0.35 M of hydrofluoric acid and 0.62 M of

2. A solution initially contains 0.50 M of sodium fluoride and 0.62 M of hydrofluoric acid. The approximate pH of 1.0 L of this solution AFTER the addition of 0.20 moles of sodium hydroxide is (Assume that the volume doesn't change, K_a for HF = 7.2×10^{-4}).
3. Calculate the ratio of $[\text{NH}_4^+]/[\text{NH}_3]$ to give a buffer with pH = 10.0.