

Neutralization

A neutralization reaction is when an **acid** and a base react to form water and a salt and involves the combination of H^+ ions and OH^- ions to generate water. The neutralization of a strong acid and strong base has a pH equal to 7. The neutralization of a strong acid and weak base will have a pH of less than 7, and conversely, the resulting pH when a strong base neutralizes a weak acid will be greater than 7.

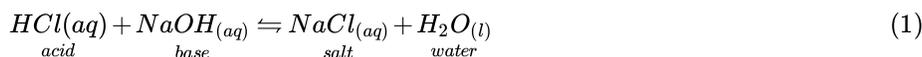
When a solution is neutralized, it means that salts are formed from equal weights of acid and base. The amount of acid needed is the amount that would give one mole of protons (H^+) and the amount of base needed is the amount that would give one mole of (OH^-). Because salts are formed from neutralization reactions with equivalent concentrations of weights of acids and bases: N parts of acid will always neutralize N parts of base.

Table 1: The most common strong acids and bases. Most everything else not in this table is considered to be weak.

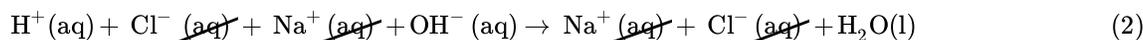
Strong Acids	Strong Bases
HCl	LiOH
HBr	NaOH
HI	KOH
HClO ₄	RbOH
HNO ₃	CsOH
	Ca(OH) ₂
	Sr(OH) ₂
	Ba(OH) ₂

Strong Acid-Strong Base Neutralization

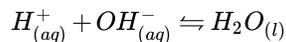
Consider the reaction between HCl and NaOH in water:



This can be written in terms of the ions (and canceled accordingly)



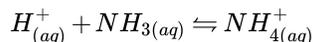
When the spectator ions are removed, the net ionic equation shows the H^+ and OH^- ions forming water in a strong acid, strong base reaction:



When a strong acid and a strong base fully neutralize, the pH is neutral. Neutral pH means that the pH is equal to 7.00 at 25 °C. At this point of neutralization, there are equal amounts of OH^- and H_3O^+ . There is no excess $NaOH$. The solution is $NaCl$ at the equivalence point. When a strong acid completely neutralizes a strong base, the pH of the salt solution will always be 7.

Weak Acid-Weak Base Neutralization

A weak acid, weak base reaction can be shown by the net ionic equation example:



The equivalence point of a neutralization reaction is when both the acid and the base in the reaction have been completely consumed and neither of them are in excess. When a strong acid neutralizes a weak base, the resulting solution's pH will be less than 7. When a strong base neutralizes a weak acid, the resulting solution's pH will be greater than 7.

Table 1: pH Levels at the Equivalence Point

Strength of Acid and Base	pH Level
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Strength of Acid and Base	pH Level
Strong Acid-Strong Base	7
Strong Acid-Weak Base	<7
Weak Acid-Strong Base	>7
Weak Acid-Weak Base	pH < 7 if $K_a > K_b$ pH = 7 if $K_a = K_b$ pH > 7 if $K_a < K_b$

Titration

One of the most common and widely used ways to complete a neutralization reaction is through [titration](#). In a titration, an acid or a base is in a flask or a beaker. We will show two examples of a titration. The first will be the titration of an acid by a base. The second will be the titration of a base by an acid.

✓ Example 1: Titrating a Weak Acid

Suppose 13.00 mL of a weak acid, with a molarity of 0.1 M, is titrated with 0.1 M NaOH. How would we draw this titration curve?

Solution

Step 1: First, we need to find out where our titration curve begins. To do this, we find the initial pH of the weak acid in the beaker before any NaOH is added. This is the point where our titration curve will start. To find the initial pH, we first need the concentration of H_3O^+ .

Set up an [ICE table](#) to find the concentration of H_3O^+ :

	HX	H_2O	H_3O^+	X^-
Initial	0.1M			
Change	-xM		+xM	+xM
Equilibrium	(0.1-x)M		+xM	+xM

$$K_a = (7)(10^{-3}) \quad (3)$$

$$K_a = (7)(10^{-3}) = \frac{(x^2)M}{(0.1 - x)M} \quad (4)$$

$$x = [H_3O^+] = 0.023 \text{ M} \quad (5)$$

Solve for pH:

$$pH = -\log_{10}[H_3O^+] = -\log_{10}(0.023) = 1.64 \quad (6)$$

Step 2: To accurately draw our titration curve, we need to calculate a data point between the starting point and the equivalence point. To do this, we solve for the pH when neutralization is 50% complete.

Solve for the moles of OH^- that is added to the beaker. We can do by first finding the volume of OH^- added to the acid at half-neutralization. 50% of 13 mL = 6.5mL

Use the volume and molarity to solve for moles (6.5 mL)(0.1M) = 0.65 mmol OH^-

Now, Solve for the moles of acid to be neutralized (10 mL)(0.1M) = 1 mmol HX

Set up an ICE table to determine the equilibrium concentrations of HX and X:

	HX	H_2O	H_3O^+	X^-

	HX	H_2O	H_3O^+	X^-
Initial	1 mmol			
Added Base		0.65 mmol		
Change	-0.65 mmol	-0.65 mmol	-0.65 mmol	
Equilibrium	0.65 mmol		0.65 mmol	

To calculate the pH at 50% neutralization, use the [Henderson-Hasselbalch approximation](#).

$$pH = pK_a + \log[\text{mmol Base}/\text{mmol Acid}]$$

$$pH = pK_a + \log[0.65\text{mmol}/0.65\text{mmol}]$$

$$pH = pK_a + \log(1)$$

$$pH = pK_a \quad (7)$$

Therefore, when the weak acid is 50% neutralized, $pH = pK_a$

Step 3: Solve for the pH at the equivalence point.

The concentration of the weak acid is half of its original concentration when neutralization is complete $0.1M/2 = .05M$ HX

Set up an ICE table to determine the concentration of OH^- :

	HX	H_2O	H_3O^+	X^-
Initial	0.05 M			
Change	-x M		+x M	+x M
Equilibrium	0.05-x M		+x M	+x M

$$K_b = (x^2)M / (0.05-x)M$$

Since $K_w = (K_a)(K_b)$, we can substitute K_w/K_a in place of K_b to get $K_w/K_a = (x^2)/(0.05)$

$$x = [OH^-] = (2.67)(10^{-7}) \quad (8)$$

$$pOH = -\log_{10}((2.67)(10^{-7})) = 6.57 \quad (9)$$

$$pH = 14 - 6.57 = 7.43 \quad (10)$$

Step 4: Solve for the pH after a bit more NaOH is added past the equivalence point. This will give us an accurate idea of where the pH levels off at the endpoint. The equivalence point is when 13 mL of NaOH is added to the weak acid. Let's find the pH after 14 mL is added.

Solve for the moles of OH^-

$$(14\text{mL})(0.1M) = 1.4\text{ mmol } OH^- \quad (11)$$

Solve for the moles of acid

$$(10\text{ mL})(0.1\text{ M}) = 1\text{ mmol } HX \quad (12)$$

Set up an ICE table to determine the OH^- concentration:

	HX	H_2O	H_3O^+	X^-
Initial	1 mmol			
Added Base		1.4 mmol		
Change	-1 mmol	-1 mmol	1 mmol	

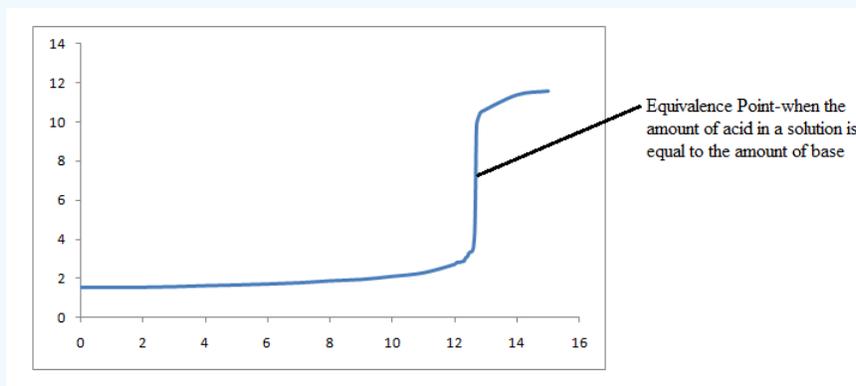
	HX	H_2O	H_3O^+	X^-
Equilibrium	0 mmol	0.4 mmol	1 mmol	

$$[OH^-] = \frac{0.4 \text{ mmol}}{10 \text{ mL} + 14 \text{ mL}} = 0.17 \text{ M} \quad (13)$$

$$pOH = -\log_{10}(0.17) = 1.8 \quad (14)$$

$$pH = 14 - 1.8 = 12.2 \quad (15)$$

We have now gathered sufficient information to construct our titration curve.



✓ Example 1

In this case, we will say that a base solution is in an Erlenmeyer flask. To neutralize this base solution, you would add an acid solution from a buret into the flask. At the beginning of the titration, before adding any acid, it is necessary to add an indicator, so that there will be a color change to signal when the equivalence point has been reached.

We can use the equivalence point to find molarity and vice versa. For example, if we know that it takes 10.5 mL of an unknown solution to neutralize 15 mL of 0.0853 M NaOH solution, we can find the molarity of the unknown solution using the following formula:

$$M_1 V_1 = M_2 V_2 \quad (16)$$

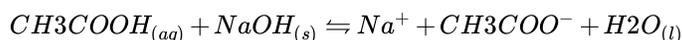
where M_1 is the molarity of the first solution, V_1 is the volume in liters of the first solution, M_2 is the molarity of the second solution, and V_2 is the volume in liters of the second solution. When we plug in the values given to us into the problem, we get an equation that looks like the following:

$$(0.0835)(0.015) = M_2(0.0105) \quad (17)$$

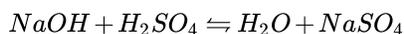
After solving for M_2 , we see that the molarity of the unknown solution is 0.119 M. From this problem, we see that in order to neutralize 15 mL of 0.0835 M NaOH solution, 10.5 mL of the .119 M unknown solution is needed.

Problems

1. Will the salt formed from the following reaction have a pH greater than, less than, or equal to seven?



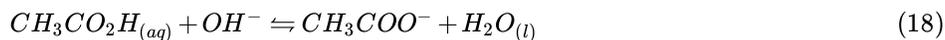
2. How many mL of .0955 M $Ba(OH)_2$ solution are required to titrate 45.00 mL of .0452 M HNO_3 ?
3. Will the pH of the salt solution formed by the following chemical reaction be greater than, less than, or equal to seven?



4. We know that it takes 31.00 mL of an unknown solution to neutralize 25.00 mL of .135 M KOH solution. What is the molarity of the unknown solution?

Solutions

1. After looking at the net ionic equation,



we see that a weak acid, CH_3CO_2H , is being neutralized by a strong base, OH^- . By looking at the chart above, we can see that when a strong base neutralizes a weak acid, the **pH level is going to be greater than 7**.

2. By plugging the numbers given in the problem in the the equation:

$$M_1V_1 = M_2V_2 \quad (19)$$

we can solve for V_2 .

$$V_2 = \frac{M_1V_1}{M_2} = \frac{(0.0452)(0.045)}{0.0955} = 21.2 \text{ mL} \quad (20)$$

Therefore it takes **21.2 mL of $Ba(OH)_2$** to titrate 45.00 mL HNO_3 .

3. We know that NaOH is a strong base and H_2SO_4 is a strong acid. Therefore, we know the **pH of the salt will be equal to 7**.

4. By plugging the numbers given in the problem into the equation:

$$M_1V_2 = M_2V_2 \quad (21)$$

we can solve for M_2 .

$$(0.135)(0.025) = M_2(0.031)$$

$M_2 = 0.108 \text{ M}$. Therefore, the molarity of the unknown solution is **.108 M**.

References

1. Petrucci, et al. General Chemistry: Principles & Modern Applications. 9th ed. Upper Saddle River, New Jersey: Pearson/Prentice Hall, 2007.
2. Criddle, Craig and Larry Gonick. The Cartoon Guide to Chemistry. New York: HarperCollins Publishers, 2005.

Contributors and Attributions

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