

Acid Base Homework 1

These questions are designed to help you go over the portion of the course that considered acids and bases. These questions are similar to those you might see on an exam.

1. What is the H^+ concentration in a 2.00 liter solution that is made with 2×10^{-7} moles HCl?

a) 1×10^7

b) 1×10^{-7}

c) $1 \times 10^{-3.5}$

d) none of these

2. What is the pH of the solution in part 1. Be as specific as you can be and give your reasons.

$pH = -\log[H^+]$ definition

$pH = -\log(1 \times 10^{-7})$

$pH = -\log(10^{-7})$

$pH = -(-7)$

$pH = 7$

I get this concentration simply by using the definition of pH, and the H^+ concentration that was obtained in problem 1.

3. If you have a 0.0500 M NaOH solution and you need 50 ml of a 0.00100 M NaOH solution, how much of the 0.0500 M solution should you use to make the dilute solution?

1. Determine how many moles of NaOH there will be in 50 mls of a 0.00100M NaOH solution.

$$50ml \left(\frac{1L}{1000ml} \right) = 0.0500L$$

$$0.0500L \text{ solution}_{(dilute)} \left(\frac{0.00100 \text{ moles NaOH}}{1L \text{ solution}_{(dilute)}} \right) = 5.00 \times 10^{-5} \text{ moles NaOH}$$

2. Determine how many liters of the concentrated solution it takes to get that many moles NaOH.

$$5.00 \times 10^{-5} \text{ moles NaOH} \left(\frac{1L \text{ solution}_{(concentrated)}}{0.0500 \text{ moles NaOH}} \right) = 1.00 \times 10^{-3} L \text{ solution}_{(concentrated)}$$

4. Would the dilute solution in question 3 be acidic or basic?

The solution would be basic. NaOH is a base. It will dissociate in water to form Na^+ and OH^- .

5. Commonly, small spills of acids or bases are cleaned up by diluting the acid or base with water. How much water would you need to add to 1ml of a 0.0500M NaOH solution in order to get a neutral solution? (A neutral solution has a $[\text{OH}^-] = 1.00 \times 10^{-7} \text{ M}$.) This is a dilution problem.

1. Find out how many moles of NaOH there are in 1ml of a 0.0500M NaOH solution. (This is your concentrated solution.)

$$1\text{ml} \left(\frac{1\text{L}}{1000\text{ml}} \right) = 0.001\text{L}$$

$$0.001\text{L solution}_{(\text{concentrated})} \left(\frac{0.0500\text{moles NaOH}}{1\text{L solution}_{(\text{concentrated})}} \right) = 5.00 \times 10^{-5} \text{moles NaOH}$$

2. What would be the final volume of a $1 \times 10^{-7} \text{ M OH}^-$ solution that held 5.00×10^{-5} moles OH^- ? (You are trying to get a solution of $1 \times 10^{-7} \text{ M OH}^-$ using a 0.05M NaOH solution.) There is 1 significant figure.

$$5 \times 10^{-5} \text{moles NaOH} \left(\frac{1\text{L solution}_{(\text{dilute})}}{1 \times 10^{-7} \text{moles NaOH}} \right) = 5 \times 10^2 \text{L solution}_{(\text{dilute})}$$

3. Since the question was how much water would be needed to make the $1 \times 10^{-7} \text{ M OH}^-$ solution, the answer is $500\text{L} - 0.001\text{L} = 499\text{L}$ water. The answer is only significant in the hundreds place to the answer is $5 \times 10^2 \text{L}$.

6. The acid H_2SO_4 has 2 protons that can be removed to form SO_4^{2-} . Write a balanced chemical equation for the reaction between H_2SO_4 and NaOH that produces SO_4^{2-} . Then write a net ionic equation for this same reaction.

$\text{NaOH} \rightarrow \text{Na}^+ + \text{OH}^-$ (NaOH is a strong base and it will dissociate completely.)

$\text{H}_2\text{SO}_4 + \text{Na}^+ + 2\text{OH}^- \rightarrow 2\text{H}_2\text{O} + \text{SO}_4^{2-} + \text{Na}^+$ (OH^- is a strong base, it accepts both H^+ from the H_2SO_4 .)

$\text{H}_2\text{SO}_4 + 2\text{OH}^- \rightarrow 2\text{H}_2\text{O} + \text{SO}_4^{2-}$ (Changing an equation to a net ionic equation just means that you should cancel out what ever appears unchanged on both sides of the equation.)

7. If 9.00×10^{-4} moles of H_2SO_4 react with NaOH, to form SO_4^{2-} , how many moles of NaOH are used?

From the chemical equations in the previous problem you can get the conversion ratios shown below.

$$9.00 \times 10^{-4} \text{ moles } H_2SO_4 \left(\frac{2 \text{ moles } OH^-}{1 \text{ mole } H_2SO_4} \right) \left(\frac{1 \text{ mole } NaOH}{1 \text{ mole } OH^-} \right) = 1.80 \times 10^{-3} \text{ moles } NaOH$$

8. If $Ba(OH)_2$ were used rather than NaOH, how many moles of $Ba(OH)_2$ would be needed?

In this case use, $Ba(OH)_2 \rightarrow Ba^{2+} + 2OH^-$

$$9.00 \times 10^{-4} \text{ moles } H_2SO_4 \left(\frac{2 \text{ moles } OH^-}{1 \text{ mole } H_2SO_4} \right) \left(\frac{1 \text{ mole } Ba(OH)_2}{2 \text{ mole } OH^-} \right) = 9.00 \times 10^{-4} \text{ moles } Ba(OH)_2$$

9. The pH scale is a measure of the acidity of a solution, it is based on the $[H^+]$ concentration of the solution. Given this why does the pH increase when the $[H^+]$ concentration goes down?

The pH of a solution is based on the $-\log[H^+]$, the negative log of the H^+ molar concentration. $\log[H^+]$ will increase as the H^+ molar concentration increases, but when $[H^+] < 1M$ these values will always be negative. For example, a 0.1M H^+ solution is more concentrated than a 0.01 M H^+ solution. When the values of $\log[H^+]$ for these solutions reflect this, $-1 > -2$. The definition of pH, however, switches the signs so the pH of the dilute solution = 2 and the pH of the more concentrated solution is 1. Thus the pH of the solution with the higher pH has the smaller $[H^+]$. (There are many ways to explain this. Examples, diagrams, etc. . Just find what you are comfortable with that makes it clear that you know how pH is defined scale and how pH relates to concentration.)

10. Which is more acidic lemon juice which has a pH of 2.2 or orange juice that has a pH of 3.6?

The lemon juice has a lower pH so it is more acidic than the orange juice.

11. What is the **approximate** $[H^+]$ concentration of each of these solutions. (You will probably find it easiest to give your answer in scientific notation.)

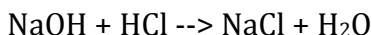
For lemon $[H^+] = 10^{-2}$ Molar, approximately

For orange juice $[H^+] = 10^{-4}$ Molar, approximately

12. You need to measure the pH of a solution precisely. You have pH paper and a pH meter available. It will take longer to use the pH meter. Which method of measuring pH should you use and why?

The question states that I need a precise measure of the pH of the solution. A pH meter is much more precise, so I would choose to use it. Even with pretty good pH paper it is difficult to get more accurate than +-0.2 pH units.

13. HCl was titrated with a 0.099 M NaOH solution. The endpoint of the titration was reached when 27.89 ml of the NaOH solution had been added. What is the number of moles of HCl in the solution. (The endpoint of the titration is when the number of moles of HCl = the number of moles of NaOH)



1. Find the number of moles NaOH used in the titration

$$27.89ml \left(\frac{1L}{1000ml} \right) = 0.02789L$$

$$0.02789L NaOH \text{ solution} \left(\frac{0.099 \text{ moles NaOH}}{1L NaOH \text{ solution}} \right) = 0.0028 \text{ moles NaOH}$$

2. Use the chemical equation for the acid base reaction to determine the number of moles HCl in the solution.

$$0.0027611 \text{ moles NaOH} \left(\frac{1 \text{ mole HCl}}{1 \text{ mole NaOH}} \right) = 0.0027611 \text{ moles HCl}$$

14. If the 30ml of a stock HCl solution was used in the titration, what is the concentration of HCl in the stock solution?

Use the definition of molar concentration to calculate the HCl concentration in the stock solution

$$30 \text{ ml} \left(\frac{1 \text{ L}}{1000 \text{ ml}} \right) = 0.030 \text{ L}$$

$$\frac{0.002761 \text{ moles HCl}}{0.030 \text{ L}} = 0.92 \text{ M HCl}$$

$$= 0.9 \text{ M HCl} \quad (\text{assuming 1 sig. fig.})$$