

pH and pOH

1. What is pH?

pH is simply a convenient way to describe the acidic character of a solution.

2. How do you solve for pH?

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

Whenever you see the “p” in front of something, it means “take the – log” of it. For example $pK_a = -\log K_a$. Remembering this can minimize confusion later.

3. What do the following values of pH indicate about sol'n acidity/basicity?

- a. pH = 7

The solution is neutral.

- b. pH > 7

The solution is basic (alkaline).

- c. pH < 7

The solution is acidic.

4. What are the rules of significant figures when solving for pH?

On the pH scale, only number after the decimal are considered significant. The digits preceding the decimal are considered place holders.

For example, the following pH has three significant figures

5. If $[H^+] = 1.0 \times 10^{-7} \text{ M}$ at 25°C , what is the pH?

$$pH = -\log(1.0 \times 10^{-7}) = 7.00$$

6. What is pOH?

pOH is simply a convenient way to describe the basic character of a solution.

7. How do you solve for pOH?

Much the same way as for pH.

$$pOH = -\log[-OH]$$

8. How do pH and pOH relate?

$$pH + pOH = 14$$

9. What do the following values of pOH indicate about sol'n acidity/basicity?

The pOH scale is the reverse of the pH scale.

- a. $pOH = 7$

The solution is neutral.

- b. $pOH > 7$

The solution is acidic.

c. $pOH < 7$

The solution is basic.

10. Calculate the pH for a 1.08M sol'n of HNO_3 .

Because HNO_3 is a strong acid, $[HNO_3] = [H^+]$

so you can simply plug the initial concentration of the strong acid in for $[H^+]$.

$$pH = -\log (1.08) = -0.333$$

A negative pH is completely reasonable. pH can go well below 0 and well above 14 – so don't get concerned if your answer isn't between 0 and 14.

11. Calculate the pH for a 1.00 M sol'n of HF. ($K_a = 7.2 \times 10^{-4}$)

Because HF is a weak acid $[HF] \gg [H^+]$. This means we cannot simply take the $-\log (1.00)$. We need to use an ICE chart to determine the $[H^+]$ produced by the solution of HF.

First, obtain the reaction equation and the expression for the equilibrium constant.



$$K_a = \frac{[H^+][F^-]}{[HF]}$$

Next, put all the data in an ICE chart.

| | $HF_{(aq)}$ | \rightleftharpoons | $H^+_{(aq)}$ | + | $F^-_{(aq)}$ |
|----------|-------------|----------------------|--------------|---|--------------|
| <i>I</i> | 1.00M | | 0 | | 0 |
| <i>C</i> | -x | | +x | | +x |
| <i>E</i> | 1.00 - x | | x | | x |

Now plug data from the equilibrium row into the expression for K_a in order to solve for $[H^+]$ at equilibrium. Remember it is the concentration of H^+ at equilibrium that will allow us to solve for the pH.

$$K_a = \frac{(x)(x)}{(1.00 - x)} = 7.2 \times 10^{-4}$$

make assumption because $K_a \ll 1$

The assumption here is that the reaction proceeds so insignificantly toward the product that the initial concentration of HF does not change enough to warrant concern.

$$K_a = \frac{x^2}{1.00} = 7.2 \times 10^{-4}$$

$$x = 2.7 \times 10^{-2}$$

Don't forget to validate the assumption. Make sure that the answer meets the 5% rule.

$$\frac{2.7 \times 10^{-2}}{1.00M} \times 100 = 2.7\%$$

(within 5% rule)

This answer is within the necessary parameters so you can solve for $[H^+]$ and then pH.

$$[H^+] = x = 2.7 \times 10^{-2} M$$

$$pH = -\log(2.7 \times 10^{-2}) = \boxed{1.569}$$

12. A 10.0 mL sample of HCl has a pH = 2.000. What volume of water must be added to change pH to 4.000?

To complete this problem you need to recognize that it is a dilution. You see that you are trying to increase the pH – which means decrease the concentration of H^+ by adding water – i.e. diluting it.

For a dilution you can use the equation:

$$M_1V_1 = M_2V_2$$

You know how much solution you started with, V_1 . You need to determine the volume of the solution after dilution, V_2 . The only things remaining are the concentration of $[H^+]$, M_1 and M_2 .

What piece of information have you been given to help you get that? The pH of the two solutions. If you can use $[H^+]$ to solve for pH – you can use pH to solve for $[H^+]$.

$$\begin{aligned} pH &= -\log [H^+] \rightarrow [H^+] = 10^{-pH} \\ [H^+]_1 &= 10^{-2.000} = 0.01000 M \\ [H^+]_2 &= 10^{-4.000} = 0.000100 M \\ M_1V_1 &= M_2V_2 \\ (0.0100M)(10.0 mL) &= (0.000100M) (V_2) \\ V_2 &= 1000 mL \end{aligned}$$

Now don't forget that this question asked for the amount of water that was added...

$$1000mL \text{ of final solution} - 10mL \text{ of original sample} = \boxed{990 mL \text{ of water added}}$$

13. Calculate the pH of a 5.0×10^{-4} M NaOH sol'n.

Because NaOH is a strong base

$$[NaOH] = [OH^-]$$

This means that you can plug concentration of NaOH into the equation for pOH.

$$pOH = -\log [OH^-] = -\log (5.0 \times 10^{-4}) = 3.30$$

Don't forget that you are solving for pH.

$$pH = 14 - pOH = 14 - 3.30 = \boxed{10.70}$$

It may seem obvious, but always make sure that you answered the question – in a hurry it is easy to forget to make sure you solved for the right thing.

14. Calculate the mass of HONH_2 required to dissolve in enough water to make 250.0 mL of sol'n having a $\text{pH} = 10.00$ ($K_b = 1.1 \times 10^{-8}$)

The first challenge to this question is identifying what HONH_2 is – strong or weak, acid or base. As this substance contains an amine group, we can recognize it is a weak base.



$$K_b = \frac{[\text{HONH}_3^+][^-\text{OH}]}{[\text{HONH}_2]}$$

| | $\text{HONH}_2(aq)$ | + | $\text{H}_2\text{O}(l)$ | \rightleftharpoons | $\text{HONH}_3^+(aq)$ | + | ^-OH |
|----------|---------------------|---|-------------------------|----------------------|-----------------------|---|---------------|
| <i>I</i> | <i>Y</i> | | — | | <i>0</i> | | <i>0</i> |
| <i>C</i> | $-x$ | | — | | $+x$ | | $+x$ |
| <i>E</i> | $Y-x$ | | — | | x | | x |

$$K_b = \frac{(x)(x)}{Y-x} = 1.1 \times 10^{-8}$$

Same initial steps – as usual. The question is asking for information linked to the value I have termed Y. If I could determine x, I would be able to solve for Y using K_b .

Is there anything in the question that would allow me to solve for x? The question gave us information about the pH which can be converted into pOH – we can use this to determine the equilibrium concentration of ^-OH which is equal to x.

$$pOH = 14 - pH = 14 - 10.00 = 4.00$$

$$[OH^-] = 10^{-pOH} = 10^{-4.00} = 1.0 \times 10^{-4} M = x$$

$$K_b = \frac{(1.0 \times 10^{-4} M)(1.0 \times 10^{-4} M)}{(Y - 1.0 \times 10^{-4} M)} = 1.1 \times 10^{-8}$$

$$Y = 0.909M$$

You know the initial concentration of the HONH₂. The question wanted the initial mass, so you will have to use dimensional analysis to convert the units.

$$0.250 \text{ L of sol'n} \frac{0.909 \text{ mol HONH}_2}{\text{L of sol'n}} \frac{33.04\text{g}}{1\text{mol HONH}_2} = \boxed{7.50\text{g of HONH}_2 \text{ required}}$$

15. The pH of a $1.00 \times 10^{-2} M$ solution of HOCN is 2.77 at 25°C. Calculate K_a of HOCN from this result.

As per the regular process of events, create your reaction equation and expression for K_a .



$$K_a = \frac{[H^+][^-OCN]}{[HOCN]}$$

Next, because we are dealing with a weak acid, we need to create an ICE chart.

| | $HOCN_{(aq)}$ | \rightleftharpoons | $H^+_{(aq)}$ | + | $^-OCN_{(aq)}$ |
|----------|---------------|----------------------|--------------|---|----------------|
| <i>I</i> | 0.0100M | | 0 | | 0 |
| <i>C</i> | -x | | +x | | +x |
| <i>E</i> | 0.0100M - x | | x | | x |

Now we are solving for the value of K_a . This means that we need the value of x. If you look at the remaining information provided in the question we have a pH value. This is actually very useful because the pH tells us the concentration of H^+ in solution at equilibrium – this value is

equivalent to x.

Remember:

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

If you rearrange this equation in terms of $[H^+]$, you get...

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

Now all you have to do is plug in values.

$$x = [H^+] = 10^{-pH} = 10^{-2.77} = 1.70 \times 10^{-3}$$

Now that the value of x has been obtained, you can readily solve for K_a .

$$K_a = \frac{[H^+][^-OCN]}{[HOCN]} = \frac{x^2}{0.0100M - x} = \frac{(1.70 \times 10^{-3})^2}{(0.0100 - 1.70 \times 10^{-3})} = \boxed{3.5 \times 10^{-4}}$$

16. Calculate the pH of a 0.30M $C_2H_5NH_2$ solution. ($K_b = 5.6 \times 10^{-4}$)

As always we want to create the reaction equation and K_b expression. Remember you can identify this substance as basic because it contains an ammonia derivative and has an overall neutral charge. So the reaction equation would be written as follows...



$$K_b = \frac{[C_2H_5NH_3^+][^-OH]}{[C_2H_5NH_2]}$$

Next step? The substance is a weak base so we need an ICE chart...

| | $C_2H_5NH_2(aq)$ | $+ H_2O(l)$ | $\rightleftharpoons C_2H_5NH_3^+(aq)$ | $+ ^-OH(aq)$ |
|----------|------------------|-------------|---------------------------------------|--------------|
| <i>I</i> | 0.30M | — | 0 | 0 |
| <i>C</i> | -x | — | +x | +x |
| <i>E</i> | 0.30 - x | — | x | x |

Plug into your expression for K_b and solve for x .

$$K_b = \frac{[C_2H_5NH_3^+][OH^-]}{[C_2H_5NH_2]} = \frac{x^2}{0.30 - x} = 5.6 \times 10^{-4}$$

Make the assumption that because K_b is small x is small enough to be inconsequential.

$$\frac{x^2}{0.30} = 5.6 \times 10^{-4}$$
$$x = 0.013M$$

Check the assumption

$$\frac{0.013}{0.30} \times 100 = 4.3\%$$

As this is under 5% the assumption is valid.

Remember that $x = [OH^-]$ this means that we can solve for pOH and use that to obtain pH.

$$pOH = -\log [OH^-] = -\log (0.013) = 1.89$$
$$pH = 14 - pOH = 14 - 1.89 = \boxed{12.11}$$