

- Solution A consists of a 1.00 M aqueous solution of HOCl at 25 °C. The pK_a of HOCl is 7.54. Calculate the pH of Solution A.

As HOCl is a weak acid, $[H^+(aq)]$ must be calculated by considering the equilibrium:

	HOCl(aq)	\rightleftharpoons	OCl ⁻ (aq)	H ⁺ (aq)
initial	1.00		0	0
change	-x		+x	+x
final	1.00 - x		x	x

The equilibrium constant K_a is given by:

$$K_a = \frac{[OCl^-(aq)][H^+(aq)]}{[HOCl]} = \frac{x^2}{(1.00-x)}$$

As $pK_a = 7.54$, $K_a = 10^{-7.54}$. K_a is very small so $1.00 - x \sim 1.00$ and hence:

$$x^2 = 1.00 \times 10^{-7.54} \quad \text{or} \quad x = 0.000170 \text{ M} = [H^+(aq)]$$

Hence, the pH is given by:

$$pH = -\log_{10}[H^+(aq)] = -\log_{10}[0.000170] = 3.77$$

$$pH = 3.77$$

At 25 °C, 1.00 L of Solution B consists of 74.4 g of NaOCl dissolved in water. Calculate the pH of Solution B.

The molar mass of NaOCl is:

$$\text{molar mass} = (22.99 \text{ (Na)} + 16.00 \text{ (O)} + 35.45 \text{ (Cl)}) \text{ g mol}^{-1} = 74.44 \text{ g mol}^{-1}$$

The number of moles present in 74.4 g is therefore:

$$\text{number of moles} = \text{mass} / \text{molar mass} = (74.4 \text{ g}) / (74.44 \text{ g mol}^{-1}) = 0.999 \text{ mol}$$

If this is present in 1.00 L, then $[OCl^-] = 0.999 \text{ M}$.

As it is a weak base, $[OH^-]$ must be calculated by considering the equilibrium:

	OCl ⁻	H ₂ O	\rightleftharpoons	HOCl	OH ⁻
initial	0.999	large		0	0
change	-y	negligible		+y	+y
final	0.999 - y	large		y	y

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The equilibrium constant K_b is given by:

$$K_b = \frac{[\text{HOCl}][\text{OH}^-]}{[\text{OCl}^-]} = \frac{y^2}{(0.999-y)}$$

For an acid and its conjugate base:

$$\text{p}K_a + \text{p}K_b = 14.00$$

$$\text{p}K_b = 14.00 - 7.54 = 6.46$$

As $\text{p}K_b = 6.46$, $K_b = 10^{-6.46}$. K_b is very small so $0.999 - y \sim 0.999$ and hence:

$$y^2 = 0.999 \times 10^{-6.46} \text{ or } y = 0.000589 \text{ M} = [\text{OH}^-]$$

Hence, the pOH is given by:

$$\text{pOH} = -\log_{10}[\text{OH}^-] = \log_{10}[0.000589] = 3.23$$

Finally, $\text{pH} + \text{pOH} = 14.00$ so

$$\text{pH} = 14.00 - 3.23 = 10.77$$

$$\text{pH} = 10.77$$

Solution B (0.40 L) is poured into Solution A (0.60 L). What amount of NaOH (in mol) must be added to give a solution, after equilibration, with a pH of 8.20?

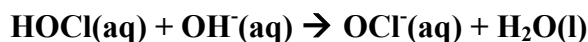
The number of moles of HOCl in 0.60 L is:

$$\begin{aligned} \text{number of moles} &= \text{concentration} \times \text{volume} \\ &= (1.00 \text{ mol L}^{-1}) \times (0.60 \text{ L}) = 0.60 \text{ mol} \end{aligned}$$

The number of moles of OCl⁻ in 0.60 L is:

$$\begin{aligned} \text{number of moles} &= \text{concentration} \times \text{volume} \\ &= (0.999 \text{ mol L}^{-1}) \times (0.40 \text{ L}) = 0.40 \text{ mol} \end{aligned}$$

The added NaOH will react with the HOCl to form more OCl⁻:



If x mol of NaOH is added then this reaction will lead to:

$$\begin{aligned} \text{number of moles of HOCl} &= (0.60 - x) \text{ mol} \\ \text{number of moles of OCl}^{\text{-}} &= (0.40 + x) \text{ mol} \end{aligned}$$

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The solution has a volume of 1.00 L so:

$$[\text{HOCl}] = (0.60 - x) \text{ M and } [\text{OCl}^-] = (0.40 + x) \text{ M}$$

Using the Henderson-Hasselbalch equation with $\text{pH} = 8.20$:

$$\text{pH} = \text{p}K_{\text{a}} + \log \frac{[\text{OCl}^-(\text{aq})]}{[\text{HOCl}(\text{aq})]} = 7.54 + \log \frac{(0.40+x)}{(0.60-x)} = 8.20$$

$$\log \frac{(0.40+x)}{(0.60-x)} = 0.66 \quad \text{or} \quad \frac{(0.40+x)}{(0.60-x)} = 10^{0.66} = 4.57$$

Solving this gives $x = 0.42 \text{ mol}$.

Answer: **0.42 mol**