

ex: The pH of blood is 7.41. Find the pOH, $[H^+]$, and $[OH^-]$.

$$pH + pOH = 14.00$$

$$7.41 + pOH = 14.00$$

$$pOH = 6.59$$



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

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

$$10^{-pOH} = 10^{\log [OH^-]}$$

$$10^{-pOH} = [OH^-]$$

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

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

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

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

"10 to the log of something equals that something."

$$10^{-7.41} = [H^+] = 3.9 \times 10^{-8} M$$

→ punch in to calculator to get ↗

$$10^{-6.59} = [OH^-] = 2.6 \times 10^{-7} M$$

* Calculating the pH of Strong Acid Solutions.

↳ HCl , HNO_3 , H_2SO_4 , $HClO_4$

- it is essential to know the major components (species) in the solution, and to know the minor species that can be ignored.

ex: $[HNO_3]$ → a very strong acid, and strong acids dissociate completely in solution.

↳ so, a 1.0M HNO_3 solution contains no HNO_3 !! Instead it contains H^+ , NO_3^- , and H_2O .

ex: Calculate the pH of a 0.10M HCl solution.



so 0.10M H^+

↳ H^+ , Cl^- , H_2O (no " HCl ")

↳ its " H^+ " contribution is negligible.
($1.0 \times 10^{-7} M H^+$)

$$pH = -\log [H^+] = -\log (0.10) \Rightarrow pH = 1.00$$

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Terry C. Jones



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