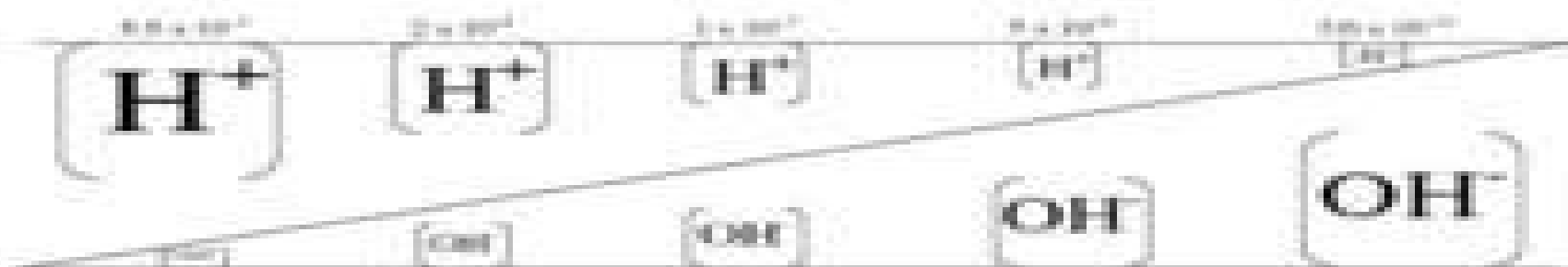


17 • Acid-Base Equilibria

IN PRACTICE



- Write the pH of each solution above the $[H^+]$ (i.e. $pH = -\log[H^+]$)
- Label the "pH" diagram as "Acidic", "Basic", and "Neutral"
- Remembering that the $[H^+]$ is $[OH^-]$ always equals 1×10^{-14} , fill in the $[OH^-]$ the each of the five solutions on the "pH" diagram. (i.e. 1×10^{-14} is called the "Autoionization Constant" for water, K_w . $[H^+] \times [OH^-] = 10^{-14}$ or $10^{-7} \times 10^{-7}$.)
- Write the "pOH" of each solution below the $[OH^-]$ (i.e.
- $pH + pOH$ always equals _____.
- A solution of acid has $[H^+] = 5.0 \times 10^{-2}$ M
 - Calculate the $[OH^-]$ _____.
 - Calculate the pH _____ the pOH _____.
- A solution of base has an $[OH^-] = 4.20 \times 10^{-2}$ M
 - Calculate the $[H^+]$ _____.
 - Calculate the pH _____ the pOH _____.
- Calculate the pH's of the following solutions:

0.01×10^{-3} M HCl pH = _____

0.01×10^{-3} M HCl pH = _____

0.01×10^{-3} M HCl pH = _____

A pH with 3 significant figures is written with _____ significant after the decimal place.

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