

- Rank $\text{H}_2\text{O}(\text{l})$, $\text{H}_2\text{S}(\text{aq})$ and $\text{HF}(\text{aq})$ in order of their Brønsted acid strengths. Explain your reasoning.

Marks
2



The H–F bond is more polar than H–O so is easier to break HF into H^+ and F^- than to break H_2O into H^+ and OH^- . Hence HF is stronger acid than H_2O .

S is much larger atom than O, so the H–S bond is much longer and weaker than H–O; hence H_2O is weaker acid than H_2S .

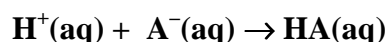
HF is stronger than H_2S because the small F^- ion bonds more strongly to the water molecules than does the larger HS^- ion.

- Buffer systems are frequently used in chemistry. Briefly describe a buffer system and how it functions? Use equations where appropriate.

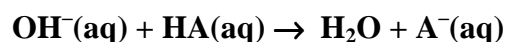
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Buffer systems resist changes in pH. They consist of a weak acid (HA) and its conjugate base (A^-) in high concentrations.

If H^+ is added, it is consumed by the A^- ions:



If OH^- is added, it is consumed by the acid present:



If the amount of H^+ or OH^- added is too great, there may be insufficient HA or A^- present and so the buffer may not work.

What ratio of concentrations of acetic acid to sodium acetate would you require to prepare a buffer with $\text{pH} = 4.00$? The K_a of acetic acid is 1.8×10^{-5} M.

The pH of a buffer can be calculated using the Henderson-Hasselbalch equation can be used:

$$\text{pH} = \text{p}K_a + \log_{10} \left(\frac{[\text{base}]}{[\text{acid}]} \right)$$

For acetic acid, $\text{p}K_a = -\log_{10}K_a = -\log_{10}(1.8 \times 10^{-5}) = 4.74$. Thus to achieve $\text{pH} 4.00$, the ratio of base to acid must be:

$$4.00 = 4.74 + \log_{10} \left(\frac{[\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]} \right)$$

$$\log_{10} \left(\frac{[\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]} \right) = -0.74 \quad \text{or} \quad \frac{[\text{CH}_3\text{COOH}]}{[\text{CH}_3\text{COO}^-]} = 5.6$$

Answer: $\text{CH}_3\text{COOH} : \text{CH}_3\text{COO}^- = 5.6$