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- Describe the difference between a strong and a weak acid.

A strong acid dissociates completely in water:

$$HA(aq) + H_2O(l) \rightarrow H_3O^+(aq) + A^-(aq)$$

As dissociation is complete, the $[H_3O^+(aq)]$ is equal to the initial concentration of HA and so the pH is given by:

$$pH = -\log(H_3O^+(aq)) = -\log([HA(aq)]_{initial})$$

A weak acid does not dissociate 100% in water:

$$HA(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + A^-(aq)$$

The position of the equilibrium and hence $[H_3O^+(aq)]$ are determined by the acid dissociation constant, K_a:

$$K_a = \frac{[H_3O^+(aq)][A^-(aq)]}{[HA(aq)]}$$

Describe in qualitative terms how the percentage ionisation of a weak acid changes when an aqueous solution of the weak acid is diluted.

Dilution increases the percentage dissocation.

Which chemical principle can be used to explain the change in percentage ionisation of a weak acid on dilution and how?

Le Chatelier's principle can be used to rationalize this effect. Increasing the amount of water shifts the equilibrium to the right.

$$HA(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + A^-(aq)$$

Dilution decreases $[H_3O^+(aq)]$ and $[A^-(aq)]$ by an equal amount and decreases $[HA(aq)]$. As $K_a = \frac{[H_3O^+(aq)][A^-(aq)]}{[HA(aq)]}$ is a constant, the decrease in $[H_3O^+(aq)]$ and $[A^-(aq)]$ must be *smaller* than that of $[HA(aq)]$ and so the percentage dissociation increases.