Saturated MCl2 Solution Equilibrium Analysis

https://gemini.google.com/app/d0374ec371ce21ec

Response: Here is the step-by-step evaluation of each option to determine the correct statement regarding a **saturated aqueous solution of** $\mathrm{MCl_2}$:

Analysis of the Equilibrium

When $\mathrm{MCl_2(s)}$ dissolves, it dissociates according to:

$$\mathrm{MCl_2(s) \rightleftharpoons M^{2+}(aq) + 2Cl^-(aq)}$$

The solubility product constant expression at a given temperature is:

$$K_{sp} = [\mathrm{M^{2+}}][\mathrm{Cl^-}]^2 = 1.0 \times 10^{-8}\ \mathrm{mol^3 dm^{-9}}$$

Because K_{sp} is a constant at constant temperature, the ion product $[\mathrm{M^{2+}}][\mathrm{Cl^-}]^2$ in a saturated solution remains fixed.

Evaluating the Options

- **(1) Evaporation of water from the solution increases $\mathrm{M^{2+}}$ and chloride ion concentrations of the solution.**
    - **Incorrect:** Evaporating water causes excess $\mathrm{MCl_2}$ to precipitate out to maintain the equilibrium concentration. The solution remains saturated, so the concentrations of $\mathrm{M^{2+}}$ and $\mathrm{Cl^-}$ remain constant.
- **(2) Chloride ion concentration of the solution can be increased by adding $\mathrm{NaCl(s)}$.**
    - **Correct:** Adding a highly soluble salt like $\mathrm{NaCl(s)}$ introduces additional $\mathrm{Cl^-}$ ions into the solution (the common ion effect). As NaCl dissolves completely, $[\mathrm{Cl^-}]$ increases. (To satisfy K_{sp}, some $\mathrm{MCl_2}$ will precipitate, thereby decreasing $[\mathrm{M^{2+}}]$ while net $[\mathrm{Cl^-}]$ increases above its initial value).
- **(3) The solution cannot be acidified by adding HCl.**
    - **Incorrect:** Adding HCl introduces $\mathrm{H^+}$ ions, which acidifies the solution.
- **(4) Chloride ion concentration of the solution cannot be increased above $1.0 \times 10^{-4}\ \mathrm{mol\ dm^{-3}}$.**
    - **Incorrect:** Let s be the molar solubility of $\mathrm{MCl_2}$:

$$K_{sp} = (s)(2s)^2 = 4s^3 = 1.0 \times 10^{-8}$$

$$s^3 = 2.5 \times 10^{-9} \implies s \approx 1.36 \times 10^{-3}\ \mathrm{mol\ dm^{-3}}$$

Therefore, the initial concentration of chloride ions is:

$$[\mathrm{Cl^-}] = 2s \approx 2.7 \times 10^{-3}\ \mathrm{mol\ dm^{-3}}$$

This value is already significantly higher than $1.0 \times 10^{-4}\ \mathrm{mol\ dm^{-3}}$, and as noted in option (2), it can be increased further by adding external sources of $\mathrm{Cl^-}$.

- **(5) Chloride ion concentration of the solution can be lowered by adding distilled water and maintaining the saturated condition.**
  - **Incorrect:** As long as the solution remains saturated (with solid $\mathrm{MCl_2}$ present), adding distilled water will simply dissolve more solid until equilibrium is re-established, keeping $[\mathrm{Cl^-}]$ constant.

Conclusion

The correct statement is **(2)**.
