

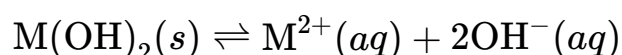
Calculating pH from Solubility Product

<https://gemini.google.com/app/67c151ac5055c765>

Response: Here is the step-by-step solution to the problem:

1. Dissolution Equilibrium and K_{sp} Expression

The dissolution equilibrium for the sparingly soluble hydroxide $M(OH)_2$ is given by:



The solubility product constant expression is:

$$K_{sp} = [M^{2+}][OH^{-}]^2$$

2. Calculate K_{sp} using the First Condition

At the initial state:

- $pH = 8.0$
- $[M^{2+}] = 1.0 \times 10^{-6} \text{ mol dm}^{-3}$

First, find the hydroxide ion concentration $[OH^{-}]$:

$$pOH = 14.0 - pH = 14.0 - 8.0 = 6.0$$

$$[OH^{-}] = 10^{-pOH} = 1.0 \times 10^{-6} \text{ mol dm}^{-3}$$

Now, substitute $[M^{2+}]$ and $[OH^{-}]$ into the K_{sp} expression:

$$K_{sp} = (1.0 \times 10^{-6}) \times (1.0 \times 10^{-6})^2 = 1.0 \times 10^{-18} \text{ mol}^3 \text{ dm}^{-9}$$

3. Calculate the New pH for the Second Condition

At the new state (same temperature, so K_{sp} remains constant):

- New $[M^{2+}] = 1.0 \times 10^{-4} \text{ mol dm}^{-3}$

Using $K_{sp} = [M^{2+}][OH^{-}]^2$:

$$1.0 \times 10^{-18} = (1.0 \times 10^{-4}) \times [OH^{-}]^2$$

$$[OH^{-}]^2 = \frac{1.0 \times 10^{-18}}{1.0 \times 10^{-4}} = 1.0 \times 10^{-14}$$

$$[OH^{-}] = 1.0 \times 10^{-7} \text{ mol dm}^{-3}$$

Next, determine pOH and pH:

$$\text{pOH} = -\log_{10}(1.0 \times 10^{-7}) = 7.0$$

$$\text{pH} = 14.0 - \text{pOH} = 14.0 - 7.0 = 7.0$$

Correct Option

Correct Option: (4) 7.0
