ANSWER KEY.

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8.11 pH and Solubility

AP Chemistry | ANSWER KEY

Show your work. Use $K_w = 1.0 \times 10^{-14}$ at 25 °C.

1. $Mg(OH)_2$ has $K_{sp} = 5.6 \times 10^{-12}$. Calculate the pH of a saturated solution of $Mg(OH)_2$ at 25 °C.
$Mg(OH)_2(s) \rightleftharpoons Mg^{2+}(aq) + 2\ OH^-(aq)$; $K_{sp} = (s)(2s)^2 = 4s^3$
$s = (5.6 \times 10^{-12} / 4)^{1/3} = 1.12 \times 10^{-4}\ M$; $[OH^-] = 2s = 2.24 \times 10^{-4}\ M$
pOH = 3.65; pH = 14.00 − 3.65 = **10.35**

2. Predict how the solubility of each salt changes when HCl is added. Then write one net ionic equation that explains your answer for CaF_2.

Salt	Solubility when HCl is added
$Mg(OH)_2$	**Increases**: H_3O^+ reacts with OH^-
AgCl	**No change**: Cl^- is not a base
CaF_2	**Increases**: $F^- + H_3O^+ \rightarrow HF + H_2O$
PbI_2	**No change**: I^- is not a base
$CaCO_3$	**Increases**: CO_3^{2-} reacts with H_3O^+

CaF_2: $\mathbf{F^-(aq) + H_3O^+(aq) \rightarrow HF(aq) + H_2O(l)}$. Removing F^- shifts $CaF_2 \rightleftharpoons Ca^{2+} + 2\ F^-$ to the right.

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3. A buffer holds the pH at 10.00. Calculate the molar solubility of $Mg(OH)_2$ in this buffer. Is it more or less soluble than in pure water (1.1×10^{-4} *M*)? Explain.

pH 10.00 gives pOH 4.00, so $[OH^-] = 1.0 \times 10^{-4}$ *M*.

$K_{sp} = [Mg^{2+}][OH^-]^2$, so $s = [Mg^{2+}] = 5.6 \times 10^{-12} / (1.0 \times 10^{-4})^2 =$ **5.6×10^{-4} *M***

More soluble. The buffer keeps $[OH^-]$ lower than in a saturated solution in pure water, so the equilibrium shifts right.

4. A solution contains 1.0×10^{-3} *M* Mg^{2+}.

(a) Will $Mg(OH)_2$ precipitate if the pH is adjusted to 10.00? Show the calculation.

(b) What is the minimum pH at which $Mg(OH)_2$ begins to precipitate?

(a) $Q = [Mg^{2+}][OH^-]^2 = (1.0 \times 10^{-3})(1.0 \times 10^{-4})^2 = 1.0 \times 10^{-11}$. Q is greater than K_{sp} (5.6×10^{-12}), so **a precipitate forms**.

(b) Set $Q = K_{sp}$: $[OH^-] = \sqrt{(5.6 \times 10^{-12} / 1.0 \times 10^{-3})} = 7.48 \times 10^{-5}$ *M*; pOH = 4.13; **pH = 9.87**

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