MULTIPLE CHOICE. Choose the one alternative that best completes the statement or answers the question.

1) Which one of the following pairs cannot be mixed together to form a buffer solution?
   A) NH₃, NH₄Cl
   B) KOH, HF
   C) H₃PO₄, KH₂PO₄
   D) NaC₂H₃O₂, HCl (C₂H₃O₂⁻ = acetate)
   E) RbOH, HBr

2) What change will be caused by addition of a small amount of HCl to a solution containing fluoride ions and hydrogen fluoride?
   A) The concentration of fluoride ion will decrease and the concentration of hydrogen fluoride will increase.
   B) The concentration of fluoride ions will increase as will the concentration of hydronium ions.
   C) The fluoride ions will precipitate out of solution as its acid salt.
   D) The concentration of hydronium ions will increase significantly.
   E) The concentration of hydrogen fluoride will decrease and the concentration of fluoride ions will increase.

3) The Henderson-Hasselbalch equation is __________.
   A) pH = pKa + log \( \frac{[base]}{[acid]} \)
   B) pH = log \( \frac{[acid]}{[base]} \)
   C) pH = pKa - log \( \frac{[base]}{[acid]} \)
   D) [H⁺] = Ka + \( \frac{[base]}{[acid]} \)
   E) pH = pKa + log \( \frac{[acid]}{[base]} \)

4) Consider a solution containing 0.100 M fluoride ions and 0.126 M hydrogen fluoride. The concentration of fluoride ions after the addition of 5.00 mL of 0.0100 M HCl to 25.0 mL of this solution is __________ M.
   A) 0.00167
   B) 0.0817
   C) 0.00253
   D) 0.0850
   E) 0.0980

5) The Ka of acetic acid is 1.7 x 10⁻⁵. The pH of a buffer prepared by combining 50.0 mL of 1.00 M potassium acetate and 50.0 mL of 1.00 M acetic acid is __________.
   A) 1.70
   B) 0.85
   C) 4.77
   D) 2.38
   E) 3.40

6) Calculate the pH of a solution prepared by dissolving 0.75 mol of NH₃ and 0.25 mol of NH₄Cl in water sufficient to yield 1.00 L of solution. The Ka of ammonia is 1.8 x 10⁻⁵.
   A) 9.73
   B) 4.27
   C) 5.22
   D) 0.89
   E) 8.78
7) Calculate the pH of a solution prepared by dissolving 0.25 mol of benzoic acid (C₇H₅O₂H) and 0.15 mol of sodium benzoate (NaC₇H₅O₂) in water sufficient to yield 1.00 L of solution. The Kₐ of benzoic acid is 6.5 x 10⁻⁵.

A) 2.39   B) 4.11   C) 3.97   D) 10.03   E) 4.19

8) The pH of a solution prepared by dissolving 0.35 mol of solid methylamine hydrochloride (CH₃NH₃Cl) in 1.00 L of 1.1 M methylamine (CH₃NH₂) is __________. The Kₐ for methylamine is 4.4 x 10⁻⁴.

A) 2.86   B) 1.66   C) 10.64   D) 11.14   E) 10.15

9) Which of the following could be added to a solution of acetic acid to prepare a buffer?

A) sodium hydroxide
B) nitric acid
C) hydrochloric acid
D) more acetic acid
E) None of the above can be added to an acetic acid solution to prepare a buffer.

10) A 25.0-mL sample of 0.723 M HClO₄ is titrated with a 0.273 M KOH solution. What is the [H⁺] (molarity) before any base is added?

A) 0.439   B) 2.81 x 10⁻¹³   C) 1.00 x 10⁻⁷   D) 0.273   E) 0.723

11) A 25.00-mL sample of 0.723 M HClO₄ is titrated with a 0.273 M KOH solution. The H₃O⁺ concentration after the addition of 10.0 mL of KOH is __________ M.

A) 0.723   B) 2.81 x 10⁻¹³   C) 0.273   D) 0.440   E) 1.00 x 10⁻⁷

12) The pH of a solution prepared by mixing 50.0 mL of 0.125 M KOH and 50.0 mL of 0.125 M HCl is __________.

A) 0.00   B) 6.29   C) 8.11   D) 5.78   E) 7.00

13) A 25.0-mL sample of an acetic acid solution is titrated with a 0.175 M NaOH solution. The equivalence point is reached when 37.5 mL of the base is added. The concentration of acetic acid is __________ M.

A) 1.83 x 10⁻⁴   B) 0.119   C) 0.263   D) 0.365   E) 0.175
14) A 25.0-mL sample of a solution of a monoprotic acid is titrated with a 0.115 M NaOH solution. The titration curve above was obtained. Which of the indicators in the table would be best for this titration?

A) bromocresol purple
B) bromthymol blue
C) thymol blue
D) phenolphthalein
E) methyl red

Consider the following table of Ksp values.

<table>
<thead>
<tr>
<th>Name</th>
<th>Formula</th>
<th>Ksp</th>
</tr>
</thead>
<tbody>
<tr>
<td>Cadmium carbonate</td>
<td>CdCO₃</td>
<td>5.2 x 10⁻¹²</td>
</tr>
<tr>
<td>Cadmium hydroxide</td>
<td>Cd(OH)₂</td>
<td>2.5 x 10⁻¹⁴</td>
</tr>
<tr>
<td>Calcium fluoride</td>
<td>CaF₂</td>
<td>3.9 x 10⁻¹¹</td>
</tr>
<tr>
<td>Silver iodide</td>
<td>AgI</td>
<td>8.3 x 10⁻¹⁷</td>
</tr>
<tr>
<td>Zinc carbonate</td>
<td>ZnCO₃</td>
<td>1.4 x 10⁻¹¹</td>
</tr>
</tbody>
</table>

15) Which compound listed below has the greatest molar solubility in water?
A) ZnCO₃      B) CdCO₃      C) CaF₂       D) AgI        E) Cd(OH)₂

16) The solubility of manganese hydroxide (Mn(OH)₂) is 2.2 x 10⁻⁵ M. What is the Ksp of Mn(OH)₂?
A) 2.1 x 10⁻¹⁴  B) 4.3 x 10⁻¹⁴  C) 1.1 x 10⁻¹⁴  D) 2.2 x 10⁻⁵  E) 4.8 x 10⁻¹⁰

17) The molar solubility of ________ is not affected by the pH of the solution.
A) AlCl₃      B) Na₃PO₄     C) NaF        D) KNO₃       E) MnS

18) In which of the following aqueous solutions would you expect AgCl to have the lowest solubility?
A) 0.020 AgNO₃
B) 0.020 M BaCl₂
C) pure water
D) 0.015 NaCl
E) 0.020 KCl
19) What is the maximum concentration (in M) of chloride ions in a solution that contains 0.100 M Pb²⁺. The $K_{sp}$ for lead (II) chloride is $1.0 \times 10^{-4}$.
   A) $1.0 \times 10^{-4}$  B) 0.058  C) 0.032  D) 0.029  E) 0.00010

20) What is the solubility (in M) of PbCl₂ in a 0.15 M solution of HCl? The $K_{sp}$ of PbCl₂ is $1.6 \times 10^{-5}$.
   A) $1.6 \times 10^{-5}$  B) $2.0 \times 10^{-3}$  C) $1.8 \times 10^{-4}$  D) $1.1 \times 10^{-4}$  E) $7.1 \times 10^{-4}$

21) For which salt should the aqueous solubility be most sensitive to pH?
   A) CaCl₂  B) Ca(NO₃)₂  C) CaBr₂  D) CaI₂  E) CaF₂

22) A solution is prepared by dissolving 0.23 mol of chloroacetic acid and 0.27 mol of sodium chloroacetate in water sufficient to yield 1.00 L of solution. The addition of 0.05 mol of HCl to this buffer solution causes the pH to drop slightly. The pH does not decrease drastically because the HCl reacts with the _______ present in the buffer solution. The $K_a$ of chloroacetic acid is $1.4 \times 10^{-3}$.
   A) chloroacetate
   B) H₃O⁺
   C) This is a buffer solution: the pH does not change upon addition of acid or base.
   D) chloroacetic acid
   E) H₂O
23) A 25.0-mL sample of a solution of an unknown compound is titrated with a 0.115 M NaOH solution. The titration curve above was obtained. The unknown compound is __________.
   a. a strong acid
   b. a weak acid
   c. a strong base
   d. a weak base
   e. neither an acid nor a base

24) A 25.0-mL sample of a solution of a monoprotic acid is titrated with a 0.115 M NaOH solution. The titration curve above was obtained. The concentration of the monoprotic acid is about __________ mol/L.

25) 25.0 mL sample of a solution of a monoprotic acid is titrated with a 0.115 M NaOH solution. The titration curve above was obtained. The $K_a$ of the monoprotic acid is about __________.

26) The pH of a solution prepared by mixing 45 mL of 0.283 M KOH and 65 mL of 0.145 M HCl is __________.

27) Which of the following could be added to a solution of potassium fluoride to prepare a buffer?
   a. hydrochloric acid
   b. nitric acid
   c. hydrofluoric acid
   d. none of these three chemicals could be used to prepare a buffer with potassium fluoride
   e. any of these three chemicals could be used to prepare a buffer with potassium fluoride
28. The addition of 0.100M AgNO₃ (aq) to a beaker containing 0.100 M NaCl (aq) results in the formation of a precipitate.

   a. What is the identity of the precipitate?
   
   b. Write the net ionic equation showing this precipitate’s formation.
   
   c. A concentrated solution of NH₃ (aq) was added to the beaker containing the precipitate. As a result, the precipitate dissolved. Write equations showing the chemical reaction which takes place when the ammonia solution is added to the beaker.

Advanced Placement Chemistry: 1994 Free Response Question

1) MgF₂(s) ⇌ Mg²⁺(aq) + 2 F⁻(aq)

In a saturated solution of MgF₂ at 18°C, the concentration of Mg²⁺ is 1.21 x 10⁻³ molar. The equilibrium is represented by the equation above.

(a) Write the expression for the solubility-product constant, K_{sp}, and calculate its value at 18°C.

(b) Calculate the equilibrium concentration of Mg²⁺ in 1.000 liter of saturated MgF₂ solution at 18°C to which 0.100 mole of solid KF has been added. The KF dissolves completely. Assume the volume change is negligible.

(c) Predict whether a precipitate of MgF₂ will form when 100.0 milliliters of a 3.00 x 10⁻³ molar Mg(NO₃)₂ solution is mixed with 200.0 milliliters of a 2.00 x 10⁻³ molar NaF solution at 18°C. Calculations to support your prediction must be shown.

(d) At 27°C the concentration of Mg²⁺ in a saturated solution of MgF₂ is 1.17 x 10⁻³ molar. Is the dissolving of MgF₂ in water an endothermic or an exothermic process? Give an explanation to support your conclusion.
MULTIPLE CHOICE. Choose the one alternative that best completes the statement or answers the question.

1) E  
   ID: chem9b 17.1-2

2) A  
   ID: chem9b 17.1-4

3) A  
   ID: chem9b 17.1-5

4) B  
   ID: chem9b 17.1-6

5) C  
   ID: chem9b 17.1-9

6) A  
   ID: chem9b 17.1-12

7) C  
   ID: chem9b 17.1-13

8) D  
   ID: chem9b 17.1-15

9) A  
   ID: chem9b 17.1-19

10) E  
   ID: chem9b 17.1-25

11) D  
   ID: chem9b 17.1-26

12) E  
   ID: chem9b 17.1-30

13) C  
   ID: chem9b 17.1-31

14) B  
   ID: chem9b 17.1-37

15) C  
   ID: chem9b 17.1-39

16) B  
   ID: chem9b 17.1-43

17) D  
   ID: chem9b 17.1-45

18) B  
   ID: chem9b 17.1-46

19) C  
   ID: chem9b 17.1-51

20) E  
   ID: chem9b 17.1-54

21) E  
   ID: chem9b 17.1-58
22) A
   ID: chem9b 17.2-9
The text content is too large to be represented here in full. Please provide a smaller snippet or consider a different format to better support your inquiry.
7. \[ pH = pK_a + \log \frac{[A^-]}{[HA]} \]
\[ = -\log(6.5 \times 10^{-5}) + \log \frac{0.15 \text{ M}}{0.25 \text{ M}} \]
\[ = 4.19 + \log 0.6 \]
\[ = 3.97 \] (C)

8. \[ pH = pK_a + \log \frac{[A^-]}{[HA]} \]
\[ pH = 10.643 + \log \left[ \frac{1.1 \text{ M}}{1.35 \text{ M}} \right] \]
\[ = 11.14 \] (D)

9. \[ \text{A, NaOH} \]
\[ \text{OH}^- + \text{H}_2\text{C}_2\text{H}_3\text{O}_2 \rightarrow \text{H}_2\text{O} + \text{C}_2\text{H}_3\text{O}_2^- \]

10. Strong acid-strong base titration, but no titration yet.
\[ [\text{strong acid}]_0 = [\text{H}^+]_{\text{equilibrium}} = 0.723 \text{ M} \] (E)

11. Strong acid-strong base titration
\[ \text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O} \]
\[ 25.00 \text{ mL} \times 0.723 \text{ mmol/L} = 18.08 \text{ mmol} \]
\[ 10.00 \text{ mL} \times 0.273 \text{ mmol/L} = 2.73 \text{ mmol} \]
\[ [\text{H}^+] = \frac{15.35 \text{ mmol}}{35.0 \text{ mL}} = 0.438 \text{ M} \] (D)

12. Strong acid-strong base at equivalence point: \( pH = 7 \)
\[ \text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O} \]
\[ 50.00 \text{ mL} \times 0.125 \text{ mmol/L} = 6.25 \text{ mmol} \]
\[ \text{but, an aqueous solution is not at equilibrium when these equal zero.} \]
\[ \text{H}_2\text{O} \rightarrow \text{H}^+ + \text{OH}^- \]
\[ K_w = 1.0 \times 10^{-14} \]
\[ X = [\text{H}^+] = [\text{OH}^-] \]
\[ x^2 = 1 \times 10^{-14} \]
\[ x = 1 \times 10^{-7} \]
13. Weak acid-strong base titration
At equivalence point, moles of acid in sample = moles of base added to sample

\[
\text{Moles of } OH^- = 0.175 \text{ moles/L} \times 0.0375 L = 0.00656 \text{ moles}
\]

\[
[\text{HC}_2\text{H}_3\text{O}_2] = \frac{0.00656 \text{ moles}}{0.0250 L} = 0.263 \text{ M}
\]

14. Looks like the equivalence point occurs at about pH=7.8 or pH=8.10
So, pick an indicator which has a pKa closest to this pH
- brom-thymine blue, pKa = 7.10

15. Well, I can directly compare CdCO\(_3\), AgI\(_3\) and ZnCO\(_3\):

\[
AB(s) \rightleftharpoons A^+ + B^-
\]

\[
K_{sp} = [A^+][B^-]
\]

\[
I\quad \text{C} \quad \text{E}
\]

\[
\text{IC} \quad +x \quad +x
\]

\[
\text{E} \quad -x \quad 0
\]

\[
\text{K}_{sp} = x^2
\]

\[
\text{K}_{sp} = \sqrt[3]{1.4 \times 10^{-11}} = 3.7 \times 10^{-4} \text{ M}
\]

So, ZnCO\(_3\) would have the biggest "x."

I can also directly compare Cd(OH)\(_2\) and CaF\(_2\):

\[
AB_2(s) \rightleftharpoons A^{2+} + 2B^-
\]

\[
K_{sp} = [A^{2+}][B^-]^2
\]

\[
\text{IC} \quad +x \quad +2x
\]

\[
\text{E} \quad -x \quad 2x
\]

\[
\text{K}_{sp} = \frac{[A^{2+}][B^-]^2}{(2x)^2} = 4x^3
\]

CaF\(_2\) will have the biggest "x."

\[
3.9 \times 10^{-11} = 4x^3
\]

\[
\left(\frac{3.9 \times 10^{-11}}{4}\right)^{\frac{1}{3}} = x = 2.1 \times 10^{-4} \text{ M}
\]

\[\text{Solubility of CaF}_2 > \text{that of ZnCO}_3\text{, so CaF}_2\text{ is most soluble.}\]
16. \( \text{Mn(OH)}_2(s) \rightleftharpoons \text{Mn}^{2+}(aq) + 2 \text{OH}^-(aq) \)

\[ K_{sp} = \frac{2.2 \times 10^{-5} \text{mole Mn(OH)}_2}{1 \text{L}} \times \frac{1 \text{mole Mn}^{2+}}{1 \text{mole Mn(OH)}_2} = 2.2 \times 10^{-5} \text{mole Mn}^{2+} \text{ at equilibrium} \]

\[ x = 2.2 \times 10^{-5} M \]

\[ K_{sp} = [\text{Mn}^{2+}][\text{OH}^-]^2 = (x)(2x)^2 = (2.2 \times 10^{-5})(2)(2.2 \times 10^{-5}) \]

\[ K_{sp} = 4.13 \times 10^{-14} \text{ } \bigcirc \]

17. (D) \( \text{KNO}_3 \) will not be affected by pH

B, C, and E contain the conjugate bases of weak acids

A is a Lewis acid

\[ \text{Al}^{3+} + 6 \text{H}_2\text{O} \rightarrow \text{Al}(	ext{OH})_6^{3+} \rightarrow \text{Al}(	ext{OH})_5\text{O}^2+ + \text{H}^+ \]

so I'm guessing that adding \( \text{H}^+(\text{lowering pH}) \) shifts equilibrium to the left, increasing \([\text{Al}^{3+}]\), so decreasing solubility of \( \text{AlCl}_3 \). Also, the soluble complex ion \( \text{Al}(	ext{OH})_4^- \) forms in basic solution. Thus a higher pH should increase solubility of \( \text{AlCl}_3 \).

18. B1 because it is 0.04M with respect to \([\text{Cl}^-]\).

The \( \text{Ag}^+ \) solutions would affect solubility, too, but not as much as 0.02M \( \text{BaCl}_2 \).

19. \( \text{PbCl}_2(s) \rightleftharpoons \text{Pb}^{2+}(aq) + 2 \text{Cl}^-(aq) \)

\[ K_{sp} = [\text{Pb}^{2+}][\text{Cl}^-]^2 \]

\[ 1.0 \times 10^{-4} = (0.100)(x)^2 \]

\[ 1 \times 10^{-3} = x^2 \]

\[ 0.032 \text{ M} = x \text{ } \bigcirc \]
(20) \[ \text{PbCl}_2(s) \rightleftharpoons \text{Pb}^{2+} + 2\text{Cl}^- \]

\[ 0 \quad 0.15 \text{ M} \]

\[ +x \quad +2x \]

\[ x \quad 0.15 + 2x \approx 0.15 \]

\[ K_{sp} = [\text{Pb}^{2+}][\text{Cl}^-]^2 \]

\[ 1.6 \times 10^{-5} = (x)(0.15)^2 \]

\[ 1.6 \times 10^{-5} \approx x = 7.1 \times 10^{-4} \text{ M} \]

\[ \frac{7.1 \times 10^{-4} \text{ mol Pb}^{2+}}{1 \text{ L}} \times \frac{1 \text{ mol PbCl}_2}{1 \text{ mol Pb}^{2+}} = 7.1 \times 10^{-4} \text{ mol PbCl}_2 \]

(21) **E** CaF\(_2\) because F\(^-\) is the conjugate base of the weak acid HF.

(22) **A** the H\(^+\) will react with the conjugate base (A\(^-\)) to form the weak acid chloroacetic acid (H\(_2\)A).

(23) **B** - the pH is higher than 7 at the equivalence point

(24) About 22 mL was added to reach equivalence point.

\[ 22 \text{ mL} \times 0.115 \text{ M} = 2.53 \text{ mmol} \]

At equivalence point, all of the acid has been consumed by base thus 2.53 mmol of acid was in sample.

\[ \frac{2.53 \text{ mmol}}{25.0 \text{ ml}} = \text{ about 0.10 M} \]
25) As a weak acid is consumed by a strong base in a titration, some of the acid is converted into its conjugate base by the reaction:

\[ HA + OH^- \rightarrow A^- + H_2O \]

Thus, between the start of the titration and the equivalence point, a buffer solution (\( A^-/HA \)) is formed. For buffer solutions,

\[ pH = pK_a + \log \left( \frac{[A^-]}{[HA]} \right) \]

At the halfway point, half of the acid has been consumed but half still remains.

\[ [HA] = [A^-] \]

Thus \( pH = pK_a + \log 1 \implies pH = pK_a \)

The equivalence point occurs at about 22 ml of added base, so the halfway point occurs at about 11 ml of added \( OH^- \).

At 11 ml of added base, the \( pH = 5.8 = pK_a \)

\[ pK_a = -\log K_a \]
\[ K_a = 10^{-pK_a} = 10^{-5.8} = 1.6 \times 10^{-6} \]

26) Strong acid - strong base reaction:

\[ OH^- + H^+ \rightarrow H_2O \]

\[
\begin{array}{c|c|c|c|c}
OH^- & H^+ & H_2O & 45 ml \times 0.283 M = 12.735 mmol & 65 ml \times 0.145 M = 9.425 mmol \\
12.735 & 9.425 & 0 & & \\
3.31 mmol & 0 & & & \\
\end{array}
\]

\[ [OH^-] = \frac{3.31 \text{ mmol}}{110 \text{ ml}} = 0.030 M \]

\[ pOH = -\log [OH^-] = -\log (0.030) = 1.52 \]

\[ pH = 14 - pOH = 14 - 1.52 = 12.48 \]
Because $F^- + H^+ \rightarrow HF$

ANY STRONG ACID WOULD DO, SO $A$ & $B$ WOULD WORK.

C (HF) IS THE CONJUGATE ACID OF $F^-$, SO THAT WOULD WORK, BE

D ANY WOULD WORK OK.

$Ag^+ + Cl^- \rightarrow AgCl(s)$ silver chloride

E $Ag^+ + 2NH_3 \rightarrow Ag(NH_3)_2^+$

thus, as this complex ion forms, the $[Ag^+]$ in the beaker decreases, which shifts the equilibrium in the reaction below to the right:

$AgCl(s) \rightleftharpoons Ag^+(aq) + Cl^-(aq)$

As the reaction shifts right to replace the missing $Ag^+$ ions en route to achieving equilibrium, the amount of $AgCl(s)$ decreases.

Thus, the addition of conc $NH_3(aq)$ to a precipitate of $AgCl(s)$ will cause the $AgCl(s)$ to dissolve.
(a) \[ K_{sp} = \left[ \text{Mg}^{2+} \right] \left[ \text{F}^- \right]^2 \]

\[ 0 + x = 1.21 \times 10^{-3} \text{M} \]
\[ x = 1.21 \times 10^{-3} \text{M} \]

\[ K_{sp} = \left[ 1.21 \times 10^{-3} \right] \left[ 2.42 \times 10^{-3} \right]^2 \]
\[ K_{sp} = 7.09 \times 10^{-9} \]

\[ [\text{KF}] = \frac{0.100 \text{mol}}{1.000 \text{L}} = 0.100 \text{M} \]
\[ [\text{F}^-] = 0.100 \text{M} \]

\[ \begin{align*}
\text{MgF}_2(s) & \longrightarrow \text{Mg}^{2+} + 2\text{F}^- \\
K_{sp} & = 7.09 \times 10^{-9}
\end{align*} \]

Since \( x \) is small, \( 0.100 + 2x \approx 0.100 \text{M} \)

\[ K_{sp} = \left[ \text{Mg}^{2+} \right] \left[ \text{F}^- \right]^2 \]

\[ 7.09 \times 10^{-9} = x (0.100)^2 \]

\[ 7.69 \times 10^{-9} \approx x \]

\[ \frac{7.69 \times 10^{-9}}{0.100^2} = x \]

\[ 7.09 \times 10^{-7} \text{M} = x = \left[ \text{Mg}^{2+} \right] \]
(c) \[
\begin{align*}
100.0 \text{ mL} & \times 3.00 \times 10^{-3} \text{ mol} = 0.300 \text{ mmol Mg(OH)}_2 \\
200.0 \text{ mL} & \times 2.00 \times 10^{-3} \text{ mol} = 0.400 \text{ mmol NaF}
\end{align*}
\]
\[
[\text{Mg}^{2+}] = \frac{0.3 \text{ mmol}}{300.0 \text{ mL}} = 0.00100 \text{ M}
\]
\[
[\text{F}^-] = \frac{0.400 \text{ mmol}}{300.0 \text{ mL}} = 0.00133 \text{ M}
\]

Reaction quotient: \(Q = [\text{Mg}^{2+}][\text{F}^-]^2 = (0.001)(0.00133)^2 = 1.78 \times 10^{-9}\)

\(Q < K\), so reaction will not shift left to produce a precipitate:
\[\text{MgF}_2(s) \rightleftharpoons \text{Mg}^{2+}(aq) + 2\text{F}^-(aq), \quad K_{sp} = 7.09 \times 10^{-9}\]

(d) At 18°C, solubility = 1.21 \times 10^{-3} M MgF₂
At 27°C, solubility of MgF₂ = 1.17 \times 10^{-3} M

As temperature of solution is increased, the solubility decreases.

\[\text{MgF}_2(s) \rightleftharpoons \text{Mg}^{2+}(aq) + 2\text{F}^-(aq) + \text{heat}\]

Therefore, the dissolving of MgF₂ must be an exothermic process, because the addition heat is causing the reaction to shift left. A shift of the left means that less MgF₂ dissolves.
1994 Free Response Answer

a. \( K_{sp} = [\text{Mg}^{2+}] \ [\text{F}^-]^2 = (1.21 \times 10^{-3})(2.42 \times 10^{-3})^2 = 7.09 \times 10^{-9} \)

b. \( x \) = change in concentration of \( \text{Mg}^{2+} \) ion  
\( 2x \) = change in concentration of \( \text{F}^- \) ion  
\( [\text{Mg}^{2+}] = (1.21 \times 10^{-3} - x) \text{ M} \)  
\( [\text{F}^-] = (0.100 + 2.42 \times 10^{-3} - 2x) \text{ M} \)

Since \( x \) is small then \((0.100 + 2.42 \times 10^{-3} - 2x) \approx 0.100 \)

\( K_{sp} = 7.09 \times 10^{-9} = (1.21 \times 10^{-3} - x)(0.100)^2 \)
\( x = 1.2092914 \times 10^{-3} \)
\( [\text{Mg}^{2+}] = 1.21 \times 10^{-3} - 1.20929 \times 10^{-3} = 7.09 \times 10^{-5} \text{ M} \)

c. \( [\text{Mg}^{2+}] = 3.00 \times 10^{-3} \text{ M} \times (100.0 \text{ mL} / 300.0 \text{ mL}) = 1.00 \times 10^{-3} \text{ M} \)
\( [\text{F}^-] = 2.00 \times 10^{-3} \text{ M} \times (200.0 \text{ mL} / 300.0 \text{ mL}) = 1.33 \times 10^{-3} \text{ M} \)

Trial \( K_{sp} = (1.00 \times 10^{-3})(1.33 \times 10^{-3})^2 = 1.79 \times 10^{-9} \)
Trial \( K_{sp} \leq 7.09 \times 10^{-9} \) results in no precipitate

d. @ 18 °C, \( 1.21 \times 10^{-3} \text{ M} \text{MgF}_2 \) dissolves  
@ 27 °C, \( 1.17 \times 10^{-3} \text{ M} \text{MgF}_2 \) dissolves

\( \text{MgF}_2 \leftrightarrow \text{Mg}^{2+} + 2\text{F}^- + \text{energy} \)

Dissolving is exothermic; if energy is increased it forces the equilibrium to shift left (according to LeChatelier’s principle) and less \( \text{MgF}_2 \) will dissolve.