Chapter 4 - Answer Key

Page 219, Quick Check

1. No. Some salts are weak electrolytes. They do not dissolve to a significant extent.
2. a) molecular
   b) ionic
   c) ionic
   d)molecular
   e) ionic
3. Sample A is tap water. It has a low electrical conductivity due to low concentrations of dissolved salts in tap water. Sample B is carbon tetrachloride. It is molecular and does not conduct electricity. Sample C is seawater. It has a significant amount of dissolved ions present so will conduct electricity well.

Page 220, Quick Check

1. $CaCl_2(aq) \rightarrow Ca^{2+}(aq) + 2Cl^-(aq)$
2. $(NH_4)_2C_2O_4(aq) \rightarrow 2NH_4^+(aq) + C_2O_4^{2-}(aq)$
3. $Na_3PO_4(aq) \rightarrow 3Na^+(aq) + PO_4^{3-}(aq)$

Page 221, Quick Check

1. a) $Na_2C_2O_4(aq) \rightarrow 2Na^+(aq) + C_2O_4^{2-}(aq)$
   0.45 M 0.90 M 0.45 M
   total [ions] = 0.90 M + 0.45 M = 1.35 M
   b) $AlCl_3(aq) \rightarrow Al^{3+}(aq) + 3Cl^-(aq)$
   2.5 M 2.5 M 7.5 M
   total [ions] = 2.5 M + 7.5 M = 10.0 M
2. $LiOH(aq) \rightarrow Li^+(aq) + OH^-(aq)$
   1.0 M 1.0 M 1.0 M
   total ion concentration = 2.0 M
   $CaCl_2(aq) \rightarrow Ca^{2+}(aq) + 2Cl^-(aq)$
   0.8 M 0.8 M 1.6 M
   total ion concentration = 2.4 M
   The $CaCl_2$ will have the greater conductivity because it has a greater ion concentration.
3. $AlBr_3(aq) \rightarrow Al^{3+}(aq) + 3Br^-(aq)$
   According to the dissociation equation, 1 mole Al^{3+} : 3 mole Br^-. If $[Br^-] = 0.15$ M, then $[Al^{3+}] = 0.050$ M.

Page 222, Practice Problems 1. a)
$[NH_4^+] = 3(3.6 \times 10^{-3}) = 1.1 \times 10^{-2}$ M
$[PO_4^{3-}] = 3.6 \times 10^{-3}$ M
b) $[H^+] = 2.6$ M $[Cl^-] = 2.6$ M
2. mass of $Al_2(SO_4)_3$ =13 g
3. total $[NH_4^+] = 0.58$ M, $[SO_4^{2-}] = 0.15$ M $[S^{2-}] = 0.14$ M

Page 223, Practice Problems

1. 5.4×10^{-2} M
2. Mass $Pb(IO_3)_2 = 7.5 \times 10^{-3}$ g
3. <u>mol</u> $CaCO_3$ 7.1×10^{-5} M

Page 224, Quick Check

1. $ZnCO_3(s) \rightleftharpoons Zn^{2+}(aq) + CO_3^{2-}(aq)$
2. $Al(OH)_3(s) \rightleftharpoons Al^{3+}(aq) + 3OH^-(aq)$
3. $Fe_2S_3(s) \rightleftharpoons 2Fe^{3+}(aq) + 3S^{2-}(aq)$

Page 226, 4.1 Review

1. a) any soluble salt or strong acid or base: NaCl, HCl, NaOH. Many answers possible.
   b) any molecular substance not including weak acids: CCl_4, SO_2, XeF_6
2. Both are acids with the same concentration, but $HClO_4$ is a strong acid that ionizes completely. H_3PO_4 is a weak acid, and remains largely in molecular form. Since there are more ions in the $HClO_4$, it will have a greater electrical conductivity.
3. a) $Mg(ClO_4)_2(aq) \rightarrow 2Mg^{2+}(aq) + 2ClO_4^-(aq)$
   b) $CaCr_2O_7(aq) \rightarrow Ca^{2+}(aq) + Cr_2O_7^{2-}(aq)$
   c) $Cu(CH_3COO)_2(aq) \rightarrow Cu^{2+}(aq) + 2CH_3COO^-(aq)$
   d) $Mn(SCN)_2(aq) \rightarrow Mn^{2+}(aq) + 2SCN^-(aq)$
   e) $Al(HC_2O_4)_3(aq) \rightarrow Al^{3+}(aq) + 3HC_2O_4^-(aq)$
   f) $Ba(OH)_2 \bullet 8H_2O(aq) \rightarrow Ba^{2+}(aq) + 2OH^-(aq) + 8\ H_2O(l)$
4. <u>mol</u> $BaSO_4 = 1.0 \times 10^{-5}$ M
5. mass $CaCO_3 = 1.8 \times 10^{-3}$ g
6. $[Na^+] = 0.025$ M $[Cr_2O_7^{2-}] = 0.013$ M
7. $[Mg^{2+}] = 0.80$ M , total $[Cl^-] = 4.9$ M, $[Na^+] = 3.3$ M
8. Measure out 1.0 L of water and add solid NaCl until there is undissolved solid remaining. Pour off until 1.0 L solution remains.
9. a$AgBrO_3(s) \rightleftharpoons Ag^+(aq) + BrO_3^-(aq)$
   b) $Al_2(CrO_4)_3(s) \rightleftharpoons 2Al^{3+}(aq) + 2CrO_4^{2-}(aq)$
   c) $Mg(OH)_2(s) \rightleftharpoons Mg^{2+}(aq) + 2OH^-(aq)$
   d) $PbSO_4(s) \rightleftharpoons Pb^{2+}(aq) + SO_4^{2-}(aq)$
   e) $Cu_3(PO_4)_2(s) \rightleftharpoons 3Cu^{2+}(aq) + 2PO_4^{3-}(aq)$

Page 230, Practice Problems

1. a) low solubility
   b) low solubility
   c) soluble
   d) soluble
   e) soluble
   d) soluble
2. (many answers possible)
   a) Na_2CO_3
   b) $SrSO_4$
   c) Ag^+
   d) NO_3^-
3. The sodium hydroxide contains sodium ions and hydroxide ions. Sodium ions form soluble compounds, so any precipitate must contain the hydroxide anion. The two compounds possible for the precipitate are $Mg(OH)_2$ and $Sr(OH)_2$. According to the Solubility Table, $Mg(OH)_2$ has low solubility and $Sr(OH)_2$ is soluble. The sample must contain strontium nitrate since no precipitate was observed.

Page 231, Quick Check

1. a) $BaSO_4$
   b) $Fe(OH)_2$
   c) $Zn_3(PO_4)_2$
   d) $CrCO_3$
   e) $Mn(OH)_2$
2. a)
   $Sr(OH)_2(aq) + 2AgNO_3(aq) \rightarrow 2AgOH(s) + Sr(NO_3)_2(aq)$
   $Sr^{2+}(aq) + 2OH^-(aq) + 2Ag^+(aq) + 2NO_3^-(aq) \rightarrow 2AgOH(s) + Sr^{2+}(aq) + 2NO_3^-(aq)$
   $Ag^+(aq) + OH^-(aq) \rightarrow AgOH(s)$

b)
   $MgS(aq) + ZnCl_2(aq) \rightarrow ZnS(s) + MgCl_2(aq)$
   $Mg^{2+}(aq) + S^{2-}(aq) + Zn^{2+}(aq) + 2Cl^-(aq) \rightarrow ZnS(s) + Mg^{2+}(aq) + 2Cl^-(aq)$
   $Zn^{2+}(aq) + S^{2-}(aq) \rightarrow ZnS(s)$

c)
   $Na_2CO_3(aq) + BaS(aq) \rightarrow BaCO_3(s) + Na_2S(aq)$
   $2Na^+(aq) + CO_3^{2-}(aq) + Ba^{2+}(aq) + S^{2-}(aq) \rightarrow BaCO_3(s) + 2Na^+(aq) + S^{2-}(aq)$
   $Ba^{2+}(aq) + CO_3^{2-}(aq) \rightarrow BaCO_3(s)$

Page 233, Practice Problems

(other answers possible)

1. First add NaCl. Filter out the $AgCl(s)$
   Then add Na_2S. Filter out the $BeS(s)$
   Finally add Na_2SO_4. Filter out the $BaSO_4(s)$
2. First add $Ca(NO_3)_2$. Filter out the $CaSO_4(s)$
   Then add $Cu(NO_3)_2$. Filter out the $CuS(s)$
   Finally add $AgNO_3$. Filter out the $AgBr(s)$
3. First add $Sr(NO_3)_2$. Filter out the $Sr_3(PO_4)_2(s)$
   Then add $Mg(NO_3)_2$. Filter out the $Mg(OH)_2(s)$
   Finally add $AgNO_3$. Filter out the $Ag_2S(s)$

Page 234, Quick Check

1. The color of the solution (for example, solutions containing Cu^{2+} are blue). Flame test color. Solubility – amount of precipitate formed, the speed at which a precipitate forms, or the ability of a precipitate to re-dissolve on the addition of NH_3 or HNO_3.
2. First add NaCl. A precipitate of both $PbCl_2$ and AgCl may form. Filter out the precipitate, dry it and measure its mass. Put the precipitate into a second beaker. To the original solution, add NaOH. If a precipitate forms, filter out the $Cu(OH)_2(s)$. To the second beaker, add 6 M NH_3. Any AgCl re-dissolves. Filter out the $PbCl_2(s)$. Dry the PbCl2(*s*) and measure its mass. If the mass had decreased, the presence of AgCl is confirmed.
3. First add HNO_3. The formation of bubbles confirms the presence of CO_3^{2-} ions. Next, add $CaCl_2$. If a precipitate forms, it is $Ca_3(PO_4)_2(s)$. Filter out the precipitate.

Page 238, 4.2 Review Questions

1. a) soluble
   b) low solubility
   c) low solubility
   d) soluble
   e) soluble
   f) soluble
2. Yes – there are silver and sulphate ions in solution. A low solubility simply means that very little will dissolve; **not** that it is insoluble.
3. mol $AgCH_3COO = 0.0665$ M
   0.0665 M < 0.1 M so $AgCH_3COO$ would have a low solubility.
4. A formula equation is written with substances as compounds and states shown.
   A complete ionic equation is written with soluble ionic compounds and strong acids in dissociated form, and compounds with low solubility undissociated. States are also shown.
   A net ionic equation only shows the substances that take part in the reaction. Spectator ions are not shown. States are also shown.
5. A spectator ion is one that does not take part in the reaction. Na^+ ions and NO_3^- ions are often spectators.
6.
   a) $(NH_4)_2S(aq) + FeSO_4(aq) \rightarrow (NH_4)_2SO_4(aq) + FeS(s)$
   $2NH_4^+(aq) + S^{2-}(aq) + Fe^{2+}(aq) + SO_4^{2-}(aq) \rightarrow 2NH_4^+(aq) + SO_4^{2-}(aq) + FeS(s)$
   $Fe^{2+}(aq) + S^{2-}(aq) \rightarrow FeS(s)$
   b) $H_2SO_3(aq) + CaCl_2(aq) \rightarrow 2HCl(aq) + CaSO_3(s)$
   $H_2SO_3(aq) + Ca^{2+}(aq) + 2Cl^-(aq) \rightarrow 2H^+(aq) + 2Cl^-(aq) + CaSO_3(s)$
   $H_2SO_3(aq) + Ca^{2+}(aq) + SO_3^{2-}(aq) \rightarrow 2H^+(aq) + CaSO_3(s)$
   c) $CuSO_4(aq) + CaS(aq) \rightarrow CuS(s) + CaSO_4(s)$
   $Cu^{2+}(aq) + SO_4^{2-}(aq) + Ca^{2+}(aq) + S^{2-}(aq) \rightarrow CuS(s) + CaSO_4(s)$
   $Cu^{2+}(aq) + SO_4^{2-}(aq) + Ca^{2+}(aq) + S^{2-}(aq) \rightarrow CuS(s) + CaSO_4(s)$
7. Na^+ and K^+ are alkali ions. Compounds containing alkali ions are soluble in water.
8. First add Na_2SO_4 to precipitate $CaSO_4$: $Ca^{2+}(aq) + SO_4^{2-}(aq) \rightarrow CaSO_4(s)$
   Filter out the precipitate
   Secondly, add Na_2S to precipitate Cr_2S_3: $2Cr^{3+}(aq) + 3S^{2-}(aq) \rightarrow Cr_2S_3\ (s)$
   Filter out the precipitate
   Finally add NaOH to precipitate $Mg(OH)_2$: $Mg^{2+}(aq) + 2OH^-(aq) \rightarrow Mg(OH)_2(s)$
   Filter out the precipitate
9. First add $Mg(NO_3)_2$ to precipitate $Mg_3(PO_4)_2$: $3Mg^{2+}(aq) + 2PO_4^{3-}(aq) \rightarrow Mg_3(PO_4)_2\ (s)$
   Filter out the precipitate
   Secondly, add $Fe(NO_3)_2$ to precipitate FeS: $Fe^{2+}(aq) + S^{2-}(aq) \rightarrow FeS(s)$
   Filter out the precipitate

Finally add $AgNO_3$ to precipitate AgCl: $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$
Filter out the precipitate

10. Nitrates are soluble in water. If a precipitate forms when a nitrate is added, the precipitate must contain the cation of the nitrate.
11. AgCl can be dissolved in 6M NH_3, while AgI cannot. This can be used to separate AgCl from AgI or AgBr. Carbonates will dissolve on the addition of a strong acid such as HNO_3.
12. a) $2Ag^+(aq) + CO_3^{2-}(aq) \rightarrow Ag_2CO_3(s)$
    b) $Ag_2CO_3(s) + 2H^+(aq) \rightarrow CO_2(g) + H_2O(l) + 2Ag^+(aq)$
    $Ag_2CO_3(s) + 4NH_3(aq) \rightarrow 2Ag(NH_3)_2^+(aq) + CO_3^{2-}(aq)$
13. a) Hard water contains Ca^{2+} ions, Fe^{2+} ions and/or Mg^{2+} ions.
    Scale is $CaCO_3$ that collects inside water heaters, pipes, kettles and boilers.
    b) $CaCO_3(s) + 2HCl(aq) \rightarrow H_2O(l) + CO_2(g) + CaCl_2(aq)$
    $CaCO_3(s) + 2H^+(aq) + 2Cl^-(aq) \rightarrow H_2O(l) + CO_2(g) + Ca^{2+}(aq) + 2Cl^-(aq)$
    $CaCO_3(s) + 2H^+(aq) \rightarrow H_2O(l) + CO_2(g) + Ca^{2+}(aq)$
    c) Many answers are possible: Na_2CO_3
    d) Water softeners remove Ca^{2+} ions and Mg^{2+} ions from hard water. Soap scum can't form without these ions.

Page 242, Quick Check

1. a) $SrCO_3(s) \rightleftharpoons Sr^{2+}(aq) + CO_3^{2-}(aq)$
   b) $Mg(OH)_2(s) \rightleftharpoons Mg^{2+}(aq) + 2OH^-(aq)$
   c) $Ca_3(PO_4)_2(s) \rightleftharpoons 3Ca^{2+}(aq) + 2PO_4^{3-}(aq)$
2. a) $K_{sp} = [Sr^{2+}][CO_3^{2-}]$
   b) $K_{sp} = [Mg^{2+}][OH^-]^2$
   c) $K_{sp} = [Ca^{2+}]^3[PO_4^{3-}]^2$
3. The solubility is the maximum *amount* of solute that can be dissolved in a particular volume of solvent. The solubility product constant is the *product* of the ion concentrations raised to the power of the coefficients from the equilibrium.

Page 244, Practice Problems

1. a) $K_{sp} = 3.7 \times 10^{-9}$
   b) $K_{sp} = 1.9 \times 10^{-13}$
   c) $K_{sp} = 1.2 \times 10^{-10}$
   d) $K_{sp} = 5.2 \times 10^{-12}$
2. $K_{sp} = 5.2 \times 10^{-12}$
3. $K_{sp} = 9.0 \times 10^{-5}$

Page 247, Practice Problems

1. a) 1.3×10^{-5} M
   b) 6.8×10^{-11} g / mL
   c) 4.5×10^{-5} M
   d) 6.8×10^{-11} g
2. 3.0×10^{-10} M
3. 4.0×10^{-3} g

Page 249, 4.3 Review Questions

1. a) $Al(OH)_3(s) \rightleftharpoons Al^{3+}(aq) + 3OH^-(aq)$ $K_{sp} = [Al^{3+}][OH^-]^3$
   b) $Cd_3(AsO_4)_2(s) \rightleftharpoons 3Cd^{2+}(aq) + 2AsO_4^{3-}(aq)$
   $K_{sp} = [Cd^{2+}]^3[AsO_4^{3-}]^2$
   c) $BaMoO_4(s) \rightleftharpoons Ba^{2+}(aq) + MoO_4^{2-}(aq)$
   $K_{sp} = [Ba^{2+}][MoO_4^{2-}]$
   d) $CaSO_4(s) \rightleftharpoons Ca^{2+}(aq) + SO_4^{2-}(aq)$
   $K_{sp} = [Ca^{2+}][SO_4^{2-}]$
   e) $Pb(IO_3)_2(s) \rightleftharpoons Pb^{2+}(aq) + 2IO_3^-(aq)$
   $K_{sp} = [Pb^{2+}][IO^{3-}]^2$
   f) $Ag_2CO_3(s) \rightleftharpoons Ag^+(aq) + CO_3^{2-}(aq)$
   $K_{sp} = [Ag^+]^2[CO_3^{2-}]$
2. a) $BaSO_3(s) \rightleftharpoons Ba^{2+}(aq) + SO_3^{2-}(aq)$
   b) The solubility is the maximum amount of solute dissolved in a given volume of solvent. The solubility product constant is the solubility squared (s^2)
3. $K_{sp} = 1.2 \times 10^{-10}$
4. $K_{sp} = 9.5 \times 10^{-17}$
5. mass gypsum (g) =0.26 g
6. a) 5.8×10^{-3} g/L   b) 7.8×10^{-3} g/L
7. $K_{sp} = 4.4 \times 10^{-36}$
8. Disagree. Even though the K_{sp} value for $CdCO_3$ (1.0×10^{-6} M) is greater than the K_{sp} for $Cd(OH)_2$ (1.2×10^{-5} M) its solubility is less.
9. The solubility of AgCl (1.3×10^{-5} M) is less than the solubility of Ag_2CrO_4 (6.5×10^{-5} M). All things being equal, the precipitate of AgCl will form first.
10. 9.0×10^{-2} g

Page 252, Practice Problems

1. TIP = 7.6×10^{-10}
   TIP > K_{sp} so a precipitate forms
2. TIP = 1.4×10^{-7}
   TIP < K_{sp} so no precipitate forms
3. TIP = 4.1×10^{-4}
   If no precipitate forms, then TIP < K_{sp} . K_{sp} must be greater than 4.1×10^{-4}.

Page 254, Practice Problems

1. a) $[Sr^{2+}] = 2.1 \times 10^{-6}$ M
   b) $[Sr^{2+}] = 2.4 \times 10^{-6}$ M
   c) $[Sr^{2+}] = 3.1 \times 10^{-5}$ M
2. 2.2 g
3. $[Ag^+] = 3.3 \times 10^{-14}$ M

Page 256, Quick Check

1. a) Only changing temperature changes the value of K_{sp}.
   b) Changing the temperature or the presence of certain ions in the solution can change the solubility of AgCl. For example, if there are Pb^{2+} ions in the solution, the solubility of AgCl would be increased.

2. To decrease the solubility, we need *more* $Mg(OH)_2$ (*s*), so we need to cause the equilibirum system to shift left. The $[Mg^{2+}]$ or $[OH^-]$ could be increased by adding $Mg(NO_3)_2$ or NaOH to the solution.
3. To increase the solubility, we need more $Mg(OH)_2$ (*s*) to *dissolve*, so we need to cause the equilibirum system to shift right. The $[Mg^{2+}]$ or $[OH^-]$ could be decreased by adding a substance to precipitate one of them out such as $Ca(NO_3)_2$ (to precipitate out the OH^-). Additionally, we could add an acid (such as HCl) to neutralize the OH^- ions.

   Since the value of K_{sp} does not change, a decrease in the $[Mg^{2+}]$ (or $[OH^-]$) causes the equilibrium above to shift right and the concentration of the other ion to increase.

Page 257 Practice Problems

1. $x = 2.7 \times 10^{-7}$ M = $[Ag^+]$ = $[AgIO_3]$ dissolved
2. $x = 8.5 \times 10^{-7}$ M = $[Pb^{2+}]$ = $[PbI_2]$ dissolved
3. 5.1×10^{-7} g/L

Page 260, 4.4 Review Questions

1. a) $FeS(s) \rightleftharpoons Fe^{2+}(aq) + S^{2-}(aq)$
   $K_{sp} = [Fe^{2+}][S^{2}]$
   b) $Mg(OH)_2(s) \rightleftharpoons Mg^{2+}(aq) + 2OH^-(aq)$
   $K_{sp} = [Mg^{2+}][OH^-]^2$
   c) $Ag_2CrO_4(s) \rightleftharpoons 2Ag^+(aq) + CrO_4^{2-}(aq)$
   $K_{sp} = [Ag^+]^2[CrO_4^{2-}]$
2. a) $CaSO_4$
   b) $CaSO_4(s) \rightleftharpoons Ca^{2+}(aq) + SO_4^{2-}(aq)$
   $K_{sp} = [Ca^{2+}][SO_4^{2-}]$

c) Not necessarily. The Ca^{2+} came from the $CaCl_2$ and the SO_4^{2-} came from the H_2SO_4. The concentrations of ions depend on the original concentrations of the $CaCl_2$ and the H_2SO_4.

3. $[Pb^{2+}] = 1.2 \times 10^{-6}$ M
4. $[C_2O_4^{2-}] = 5 \times 10^{-7}$ M
5. $[CO_3^{2-}] = 16$ M
6. 3×10^{-6} g
7. TIP = $[Sr^{2+}][SO_4^{2-}] = (9.2 \times 10^{-5})(1.1 \times 10^{-2}) = 1.0 \times 10^{-6}$
   TIP > K_{sp} so precipitate forms
8. TIP = $[Pb^{2+}][Cl^-]^2 = (1.0 \times 10^{-4})(0.080)^2 = 6.6 \times 10^{-7}$ TIP < K_{sp} so no precipitate forms
9. a) $K_{sp} = [Ag^+][Cl^-] = 1.8 \times 10^{-10}$
   $K_{sp} = [Ag^+][I^-] = 8.5 \times 10^{-17}$

b) See above

c) AgI

d) $[Ag^+] = 4.3 \times 10^{-15}$ M

e) $[Ag^+] = 9.0 \times 10^{-9}$ M

f) $4.3 \times 10^{-15} \text{ M} < [Ag^+] < 9.0 \times 10^{-9}$ M

g) $[I^-] = 9.4 \times 10^{-9}$ M

h) % I^- remaining = $\dfrac{9.4 \times 10^{-9}\text{ M}}{0.020\text{ M}} \times 100\% = 0.00047\ \%$

so 100.00000 – 0.00047 = 99.999953 % precipitated out.

10. 4.0 g
11. To decrease the solubility, the equilibrium must be shifted left. By adding $Pb(NO_3)_2$ or KI, the common ions (Pb^{2+} or I^-) cause an increase in their respective concentrations.
12. In Na_2SO_4, the common ion SO_4^{2-} exists, so equilibrium shifts left and solubility is decreased. In water, there are no Ba^{2+} or SO_4^{2-} ions.
13. In HCl, the H^+ ions will neutralize the OH^- ions, causing the $[OH^-]$ to decrease. The equilibrium will shift right and the solubility of $Fe(OH)_3$ will increase. It is more soluble in HCl than in water.
14. a) 4.7×10^{-4} g
    b) 5.1×10^{-8} g