1.	solubility = 5.59 $\frac{g}{L} \times \frac{1 \text{ mol}}{84.0 \text{ g}} = 6.65 \times 10^{-2} \text{ M}$
2.	solubility = $\frac{0.99 \text{ g}}{0.100 \text{ L}} \times \frac{1 \text{ mol}}{278.2 \text{ g}} = 3.6 \times 10^{-2} \text{ M}$
3.	solubility = $1.26 \times 10^{-3} \frac{\text{mol}}{\text{L}} \times \frac{84.3 \text{ g}}{1 \text{ mol}} = 0.106 \text{ g/L}$
4.	solubility = $1.2 \times 10^{-4} \frac{\text{mol}}{\text{L}} \times \frac{275.8 \text{ g}}{1 \text{ mol}} \times 0.100 \text{ L} = 3.3 \times 10^{-3} \text{ g}$ (in 100 mL)
5.	solubility = $\frac{92.6 \text{ g}}{0.150 \text{ L}} \times \frac{1 \text{ mol}}{100.0 \text{ g}}$ = 6.17 M
6.	mass AgClO ₂ = 0.014 $\frac{\text{mol}}{\text{L}} \times 0.0500 \text{ L x} \frac{175.4 \text{ g}}{1 \text{ mol}} = 0.12 \text{ g}$
7.	mass MnCl ₂ = 5.75 $\frac{\text{mol}}{\text{L}} \times 0.125 \text{ L} \times \frac{125.9 \text{ g}}{1 \text{ mol}}$ = 90.5 g
8.	 (a) 84.84 g - 54.87 g = 29.97 g (b) 62.59 g - 54.87 g = 7.72 g (c) Since "solution" = water + solid, then water = solution - solid mass of water = 29.97 g - 7.72 g = 22.25 g
	(d) $\frac{7.72 \text{g}\text{KCI}}{22.25 \text{g}\text{water}} = \frac{X \text{g}\text{KCI}}{100.0 \text{g}\text{water}}$, and $X = 34.7 \text{g}$
	(e) solubility = $\frac{7.72 \text{ g}}{0.02500 \text{ L}} \times \frac{1 \text{ mol}}{74.6 \text{ g}} = 4.14 \text{ M}$
9.	(a) $147.42 \text{ g} - 87.23 \text{ g} = 60.19 \text{ g}$ of solution (b) $104.08 \text{ g} - 87.23 \text{ g} = 16.85 \text{ g}$ of $(NH_4)_2SO_4$ (c) $60.19 \text{ g} - 16.85 \text{ g} = 43.34 \text{ g}$ of water
	(d) $\frac{16.85 \text{g} (\text{NH}_4)_2 \text{SO}_4}{43.34 \text{g water}} = \frac{X \text{g} (\text{NH}_4)_2 \text{SO}_4}{100.0 \text{g water}}$, and $X = 38.88 \text{g}$
	(e) $[(NH_4)_2SO_4] = \frac{16.85 \text{ g}}{0.0700 \text{ L}} \times \frac{1 \text{ mol}}{132.1 \text{ g}} = 1.82 \text{ M}$
10.	(a) low solubility(c) low solubility(e) low solubility(g) soluble(i) soluble(b) soluble(d) soluble(f) soluble(h) low solubility(j) low solubility
11.	 (a) AgBr is a precipitate (b) no precipitate (c) Al(OH)₃ is a precipitate (d) Pbl₂ is a precipitate (e) BaSO₄ is a precipitate (f) no precipitate
12.	(a) BaSO ₄ has a very low solubility, so that the [Ba ²⁺] is very low and hence there is little toxicity. (What can't dissolve, can't poison you.)
	(b) AgBr also has a very low solubility: a saturated solution at 18°C has a silver ion concentration of 69 parts per billion (69 g per 10 ⁶ L of water). Ag ⁺ ion is toxic to micro–organisms at concentrations as low as 10 parts per billion, yet is harmless to humans at such low concentrations.

13. (a) $Pb(NO_3)_2$ and NaCl (b) $AgNO_3$ and NaBr (c) $Cr(NO_3)_3$ and Na_2S (d) $Sr(NO_3)_2$ and Na_2SO_4

- 14. (a) $MgS(aq) + Sr(OH)_2(aq) \longrightarrow Mg(OH)_2(s) + SrS(aq); Mg(OH)_2(s)$ has a low solubility.
 - (b) $CuBr_2(aq) + Pb(NO_3)_2(aq) \longrightarrow PbBr_2(s) + Cu(NO_3)_2(aq)$; $PbBr_2(s)$ has a low solubility.
 - (c) all products soluble
 - (d) $Ba(NO_3)_2(aq) + Li_2SO_4(aq) \longrightarrow BaSO_4(s) + 2 LiNO_3(aq); BaSO_4(s)$ has a low solubility.
 - (e) $2 \text{ K}_3 \text{PO}_4(\text{aq}) + 3 \text{ CuCl}_2(\text{aq}) \longrightarrow \text{Cu}_3(\text{PO}_4)_2(\text{s}) + 6 \text{ KCl}(\text{aq}); \text{ Cu}_3(\text{PO}_4)_2(\text{s}) \text{ has a low solubility.}$
 - (f) $3 (NH_4)_2SO_3(aq) + Al_2(SO_4)_3(aq) \longrightarrow Al_2(SO_3)_3(s) + 3 (NH_4)_2SO_4(aq);$

 $Al_2(SO_3)_3(s)$ has a low solubility.

- (g) $3 \text{ AgNO}_3(aq) + \text{Na}_3\text{PO}_4(aq) \longrightarrow \text{Ag}_3\text{PO}_4(s) + 3 \text{ NaNO}_3(aq); \text{Ag}_3\text{PO}_4(s) \text{ has a low solubility.}$
- (h) all products soluble
- (i) $CoSO_4(aq) + Li_2CO_3(aq) \longrightarrow CoCO_3(s) + Li_2SO_4(aq); CoCO_3(s)$ has a low solubility.
- (j) $2 \operatorname{Fe}(\operatorname{NO}_3)_3(\operatorname{aq}) + 3 \operatorname{MgS}(\operatorname{aq}) \longrightarrow \operatorname{Fe}_2 S_3(s) + 3 \operatorname{Mg}(\operatorname{NO}_3)_2(\operatorname{aq}); \operatorname{Fe}_2 S_3(s)$ has a low solubility.
- (k) $BeSO_4(aq) + (NH_4)_2CO_3(aq) \longrightarrow BeCO_3(s) + (NH_4)_2SO_4(aq); BeCO_3(s)$ has a low solubility.
- (I) MgSO₄(aq) + Sr(OH)₂(aq) \longrightarrow Mg(OH)₂(s) + SrSO₄(s); both products have low solubility.
- 15. The salt must be **soluble** so as to obtain a sufficiently high concentration of the anion to be added.
- 16. (a) Ag^+ , Ba^{2+} , Pb^{2+} or Ca^{2+} (b) Sr^{2+}
- You cannot differentiate between Ag⁺ and Pb²⁺. Any anion that will precipitate Ag⁺ will also precipitate Pb²⁺, and vice versa. (However, other information such as colour of the precipitate can eventually lead to differentiation.)

	Cl	SO ₄ ²⁻	S ^{2–}	OH_	PO ₄ ^{3–}
AI ³⁺		_	ppt	ppt	ppt
Ag^+	ppt	ppt	ppt	ppt	ppt

18. Analyze for solubility.

Precipitate the Ag^+ with NaCl or Na₂SO₄. Then precipitate the Al^{3+} with Na₂S, NaOH or Na₃PO₄.

19. Analyze for solubility.

	Cl	SO ₄ ²⁻	S ^{2–}	OH	PO ₄ ^{3–}
Sr ²⁺		ppt			ppt
Ca ²⁺		ppt		ppt	ppt
Ag ⁺	ppt	ppt	ppt	ppt	ppt

First precipitate the Ag^{\dagger} by adding NaCl or Na₂S.

Then precipitate the Ca^{2+} by adding NaOH.

Finally, precipitate the Sr^{2+} by adding Na₂SO₄ or Na₃PO₄.

20. Analyze for solubility.

	CI	SO ₄ ²⁻	S ^{2–}	OH_	PO ₄ ^{3–}
Mg ²⁺		_		ppt	ppt
Pb ²⁺	ppt	ppt	ppt	ppt	ppt
Zn ²⁺		_	ppt	ppt	ppt

First precipitate the Pb^{2+} by adding Na₂SO₄ or NaCl. Then precipitate the Zn²⁺ by adding Na₂S. Finally, precipitate the Mq^{2+} by adding NaOH or Na₃PO₄.

21. Analyze for solubility.

-	Cl	SO ₄ ²⁻	S ^{2–}	OH_	PO ₄ ^{3–}
Fe ³⁺			ppt	ppt	ppt
Ca ²⁺	_	ppt	_	ppt	ppt
Ag ⁺	ppt	ppt	ppt	ppt	ppt
Be ²⁺	_	_	_	ppt	ppt

First precipitate the Ag^+ by adding NaCl.

Next, **either** precipitate Ca^{2+} by adding Na₂SO₄, followed by the addition of Na₂S to precipitate Fe³⁺ and finally either NaOH or Na₃PO₄ to precipitate Be²⁺.

or precipitate Fe^{3+} by adding Na₂S, followed by the addition of Na₂SO₄ to precipitate Ca²⁺ and finally either NaOH or Na₃PO₄ to precipitate Be^{2+} .

22. **Step 1**: Take 1 mL of the solution which might contain Ag^+ and/or Al^{3+} . Add a few drops of 1 M NaCl solution.

If there is NO PRECIPITATE then Ag^+ is absent. Proceed to Step 2. If a PRECIPITATE FORMS then Ag^+ is present. Filter off and discard the precipitate and proceed to Step 2 to test the rest of the solution.

Step 2: To the solution from Step 1, add a few drops of 1 M Na₃PO₄ solution.

If there is NO PRECIPITATE then Al³⁺ is absent. If a PRECIPITATE FORMS then Al³⁺ is present.

23. Step 1: Take 1 mL of the solution which might contain Ag^+ , Ca^{2+} and/or Sr^{2+} . Add a few drops of 1 M NaCl solution. If there is NO PRECIPITATE then **Ag⁺ is absent**. Proceed to Step 2.

If a PRECIPITATE FORMS then Ag⁺ is present. Filter off and discard the precipitate and proceed to Step 2 to test the rest of the solution.

- **Step 2:** To the solution from Step 1, add a few drops of 1 M NaOH solution. If there is NO PRECIPITATE then Ca^{2^+} is absent. Proceed to Step 3. If a PRECIPITATE FORMS then Ca^{2+} is present. Filter off and discard the precipitate and proceed to Step 3 to test the rest of the solution.
- Step 3: To the solution from Step 2, add a few drops of 1 M Na₂SO₄. If there is NO PRECIPITATE then Sr²⁺ is absent. If a PRECIPITATE FORMS then **Sr²⁺ is present**.

24. (a) Qualitative analysis

(b) Analyze for solubility.

	Ag^+	Ca ²⁺	Mg ²⁺
Ē	ppt	_	_
SO ₄ ²⁻	ppt	ppt	
OH⁻	ppt	ppt	ppt

Add $Mg(NO_3)_2$ first to precipitate any OH⁻; any other cation would precipitate more than one anion. If a precipitate occurs then OH⁻ is present, otherwise OH⁻ is absent.

- (c) After filtering off any precipitate, add Ca(NO₃)₂. If a precipitate occurs then SO₄²⁻ is present; if not, SO₄²⁻ is absent. Filter off and discard any precipitate and add AgNO₃. If a precipitate occurs then I⁻ is present; if not, I⁻ is absent.
- 25. Analyze for solubility.

-	Ag ⁺	Cu ²⁺	Ba ²⁺	Sr ²⁺
S ^{2–}	ppt	ppt	_	_
OH_	ppt	ppt	ppt	
CI	ppt	_	_	
CO ₃ ²⁻	ppt	ppt	ppt	ppt

- **Step 1:** Add $Sr(NO_3)_2$ to the test solution. If a precipitate forms, CO_3^{2-} is present. Filter off and discard any precipitate and proceed to Step 2.
- **Step 2:** Add Ba(NO₃)₂ to the test solution. If a precipitate forms, OH⁻ is present. Filter off and discard any precipitate and proceed to Step 3.
- **Step 3:** Add Cu(NO₃)₂ to the test solution. If a precipitate forms, S²⁻ is present. Filter off and discard any precipitate and proceed to Step 4.
- **Step 4:** Add AgNO₃ to the test solution. If a precipitate forms, Cl⁻ is present.

26. (a) mass BaSO₄ = 3.75 g - 1.21 g = 2.54 g; moles BaSO₄ = 2.54 g x
$$\frac{1 \text{mol}}{233.4 \text{ g}}$$
 = 0.0109 mol
(b) moles Ba²⁺ in solution = moles BaSO₄ precipitated ; [Ba²⁺] = $\frac{0.0109 \text{ mol}}{0.0250 \text{ L}}$ = 0.435 M

27. moles Pb^{2^+} = moles $PbSO_4$ = 4.28 g x $\frac{1 \text{mol}}{303.3 \text{ g}}$ = 0.0141 mol ; $[Pb^{2^+}] = \frac{0.0141 \text{mol}}{0.1000 \text{ L}}$ = 0.141 M

28. moles
$$Ag^+ = 1.57 \text{ g } Ag_3PO_4 \times \frac{1 \text{mol } Ag_3PO_4}{418.7 \text{ g } Ag_3PO_4} \times \frac{3 \text{ mol } Ag^+}{1 \text{ mol } Ag_3PO_4} = 0.0112 \text{ mol}$$

 $[Ag^+] = \frac{0.0112 \text{ mol}}{0.02500 \text{ L}} = 0.450 \text{ M}$

29. First, find the moles of I_2 titrated with the Na₂S₂O₃ :

moles $I_2 = 3.61 \times 10^{-3} L \times \frac{0.0122 \text{ mol Na}_2 S_2 O_3}{L} \times \frac{1 \text{ mol } I_2}{2 \text{ mol Na}_2 S_2 O_3} = 2.202 \times 10^{-5} \text{ mol}$

Next, use the combined reaction equation to find the moles of O₂ in the original sample:

moles
$$O_2 = 2.202 \times 10^{-5} \text{ mol } I_2 \times \frac{0.5 \text{ mol } O_2}{1 \text{ mol } I_2} = 1.101 \times 10^{-5} \text{ mol}$$

Finally, calculate the concentration of O₂, in milligrams per litre, in the original sample:

concentration
$$O_2 = \frac{1.101 \times 10^{-5} \text{ mol}}{150.0 \times 10^{-3} \text{ L}} \times \frac{32.0 \text{ g}}{1 \text{ mol}} \times \frac{1 \text{ mg}}{10^{-3} \text{ g}} = \frac{2.35 \text{ mg}}{\text{L}}$$

30. First, find the moles of I_2 titrated with the Na₂S₂O₃ :

moles
$$I_2 = 11.57 \times 10^{-3} L \times \frac{0.00948 \text{ mol Na}_2 S_2 O_3}{L} \times \frac{1 \text{ mol } I_2}{2 \text{ mol Na}_2 S_2 O_3} = 5.484 \times 10^{-5} \text{ mol}$$

Next, use the combined reaction equation to find the moles of O₂ in the original sample:

moles
$$O_2 = 5.484 \times 10^{-5} \text{ mol } I_2 \times \frac{0.5 \text{ mol } O_2}{1 \text{ mol } I_2} = 2.742 \times 10^{-5} \text{ mol}$$

Finally, calculate the concentration of O₂, in milligrams per litre, in the original sample:

concentration
$$O_2 = \frac{2.742 \times 10^{-5} \text{ mol}}{300.0 \times 10^{-3} \text{ L}} \times \frac{32.0 \text{ g}}{1 \text{ mol}} \times \frac{1 \text{ mg}}{10^{-3} \text{ g}} = \frac{2.92 \text{ mg}}{\text{L}}$$

- 31. [H⁺] = 0.10 M, [OH⁻] = 1.0 x 10⁻¹³ M, pH = 1
- 32. When pH decreases, [H⁺] increases and [OH⁻] decreases.
- 33. Black coffee has a higher pH than stomach acid and therefore a higher [OH⁻].
- 34. Dissolving CO_2 in water increases the $[H^+]$ and lowers the pH below 7.
- 35. (a) Increasing the pH by 2 units decreases the $[H^+]$ by 2 powers of 10, and $[H^+]$ is 100 times less.
 - (b) Decreasing the pH by 4 units increases the $[H^+]$ by 4 powers of 10 and $[H^+]$ is 10,000 times more.
 - (c) Decreasing the pH by 7 units increases the $[H^+]$ by 7 powers of 10, and $[H^+]$ is 10⁷ times more.
 - (d) Increasing the pH by 3 units decreases the [H⁺] by 3 powers of 10, and [H⁺] is 1000 times less.
- 36. pH = 5
- 37. Solution A has a neutral pH, solution B has a basic pH and solution C has an acidic pH. Solution B has the highest pH and solution C has the lowest pH.
- 38. total moles HCl added = $25.0 \times 10^{-3} L \times 0.153 \text{ mol/L} = 0.003825 \text{ mol}$

moles NaOH = 12.4×10^{-3} L x 0.200 mol/L = 0.00248 mol = excess moles HCl

moles HCl reacted by NH₃ = total moles HCl added – excess moles HCl = 0.003825 mol – 0.00248 mol = 0.001345 mol = moles NH₃ reacted

but: moles NO_3^- = moles NH_3 = 0.001345 mol

percentage nitrate in sample = $\frac{0.001345 \text{ mol NO}_3^-}{1.25 \text{ g sample}} \times \frac{62.0 \text{ g NO}_3^-}{\text{mol NO}_3^-} \times 100 \% = 6.67 \%$

39. total moles HCl added = $100.0 \times 10^{-3} L \times 0.2006 \text{ mol/L} = 0.02006 \text{ mol}$ moles NaOH = $28.4 \times 10^{-3} L \times 0.199 \text{ mol/L} = 0.00565 \text{ mol} = excess \text{ moles HCl}$ moles HCl reacted by NH₃ = total moles HCl added – excess moles HCl = $0.02006 \text{ mol} - 0.00565 \text{ mol} = 0.01441 \text{ mol} = \text{moles NH}_3 \text{ reacted}$

but: moles NO_3^- = moles NH_3 = 0.01441 mol

percentage nitrate in sample = $\frac{0.01441 \text{ mol NO}_3^-}{1.02 \text{ g sample}} \times \frac{62.0 \text{ g NO}_3^-}{\text{mol NO}_3^-} \times 100 \% = 87.6 \%$

40. # of moles/L = $\frac{0.150 \times 10^{-3} \text{ g}}{\text{L}} \times \frac{1 \text{ mol}}{62.0 \text{ g}} = 2.42 \times 10^{-6} \text{ M}$



According to the graph, the concentration of nitrate in the spectrophotometer sample is 9.0 mg/L. Since the original water sample was diluted from 50.0 mL to 250.0 mL, then:

Original concentration of nitrate = $\frac{9.0 \text{ mg}}{\text{L} \text{ (diluted)}} \times \frac{250.0 \text{ mL} \text{ (diluted)}}{50.0 \text{ mL} \text{ (original)}} = 45 \text{ mg/L}$

42. mass of P = 1.61 g NH₄MgPO₄•6H₂O x $\frac{1 \text{mol} \text{NH}_4 \text{MgPO}_4 \bullet 6\text{H}_2 \text{O}}{245.3 \text{ g NH}_4 \text{MgPO}_4 \bullet 6\text{H}_2 \text{O}} \times \frac{1 \text{mol} \text{P}_4 \text{MgPO}_4 \bullet 6\text{H}_2 \text{O}}{1 \text{mol} \text{NH}_4 \text{MgPO}_4 \bullet 6\text{H}_2 \text{O}}$

x
$$\frac{31.0 \text{ gP}}{1 \text{ mol P}} = 0.2035 \text{ g}$$

percentage P in sample = $\frac{\text{mass P in product}}{\text{mass of sample}} \times 100 \% = \frac{0.2035 \text{ g}}{2.00 \text{ g}} \times 100 \% = 10.2 \%$

This result is NOT within 1 % of the claim made by Brand X.

43. mass of P = 1.55 g NH₄MgPO₄•6H₂O x
$$\frac{1 \text{mol NH}_4 \text{MgPO}_4 \cdot 6\text{H}_2 \text{O}}{245.3 \text{ g NH}_4 \text{MgPO}_4 \cdot 6\text{H}_2 \text{O}} \times \frac{1 \text{mol P}}{1 \text{mol NH}_4 \text{MgPO}_4 \cdot 6\text{H}_2 \text{O}} \times \frac{31.0 \text{ g P}}{1 \text{mol P}} = 0.1959 \text{ g}$$

mass P in product 0.1959 g

percentage P in sample =
$$\frac{\text{mass P in product}}{\text{mass of sample}} \times 100 \% = \frac{0.1959 \text{ g}}{5.27 \text{ g}} \times 100 \% = 3.72 \%$$



According to the graph, the concentration of phosphate in the spectrophotometer sample is 1.13 mg/L. Since the original water sample was diluted from 25.0 mL to 100.0 mL, then:

Original concentration of phosphate = $\frac{1.13 \text{ mg}}{\text{L}(\text{diluted})} \times \frac{100.0 \text{ mL}(\text{diluted})}{25.0 \text{ mL}(\text{original})} = 4.52 \text{ mg/L}$