11. Calculate the molar solubility of calcium oxalate (CaC_2O_4). Show all of your steps clearly.

$CaC_2O_{4(s)} \rightarrow Ca^{2+} + C_2O_4^{2-}$

1s   1s   1s

$K_{sp} = [Ca^{2+}][C_2O_4^{2-}] = 2.3 \times 10^{-9}$ ← look up in solubility table

$2.3 \times 10^{-9} = (s)(s) = s^2$

$s = \sqrt{2.3 \times 10^{-9}} = \boxed{4.8 \times 10^{-5} M}$

(3 s.f.) 12. If 250.0 mL of 0.000340M $Cu(NO_3)_2$ is mixed with 350.0 mL of 3.12×10^{-4}M KIO_3, will a precipitate form? Show all of your steps clearly.

balanced double-repl.: $Cu(NO_3)_2 + 2KIO_3 \rightleftharpoons Cu(IO_3)_{2(s)} + 2KNO_3$ · possible ppt is $Cu(IO_3)_2(s)$

net ionic: $Cu(IO_3)_{2(s)} \rightarrow Cu^{2+} + 2IO_3^-$

$[Cu^{2+}]_{dil} = \frac{C_1V_1}{V_2} = \frac{(3.4 \times 10^{-4} M)(250.0 mL)}{600.0 mL} = 1.417 \times 10^{-4} M$

$[IO_3^{2-}]_{dil} = \frac{C_1V_1}{V_2} = \frac{(3.12 \times 10^{-4})(350.0 mL)}{600.0 mL} = 1.82 \times 10^{-4} M$

$TIP = [Cu^{2+}]_{dil}[IO_3^{2-}]_{dil} = (1.417 \times 10^{-4})(1.82 \times 10^{-4})^2 = 4.69 \times 10^{-12} < 6.9 \times 10^{-8}$

TIP < Ksp

∴ no ppt

(2 s.f) 13. Calculate the maximum [F⁻] that can exist in a solution in which $[Sr^{2+}] = 0.00050$ M. Show all of your steps clearly.

$5.0 \times 10^{-4} M$

$SrF_{2(s)} \rightarrow Sr^{2+} + 2F^-$

$K_{sp} = [Sr^{2+}][F^-]^2 = 4.3 \times 10^{-9}$

$[F^-]^2 = \frac{K_{sp}}{[Sr^{2+}]}$

$[F^-] = \sqrt{\frac{K_{sp}}{[Sr^{2+}]}} = \sqrt{\frac{4.3 \times 10^{-9}}{5.0 \times 10^{-4}}} = \boxed{2.9 \times 10^{-3} M}$

(2 s.f.) 14. Calculate the mass of Na_2CO_3 that must be added to 2.50 L of 0.00085 M $MgCl_2$ in order to just start precipitation. Show all of your steps clearly. -look up Ksp in solubility table.

↳ TIP = Ksp

$Na_2CO_3 + MgCl_2 \rightleftharpoons MgCO_{3(s)} + 2NaCl$

↑ possible ppt.

net ionic: $MgCO_3 \rightarrow Mg^{2+} + CO_3^{2-}$

$Ksp = [Mg^{2+}][CO_3^{2-}] = TIP$

$[CO_3^{2-}] = \frac{Ksp}{[Mg^{2+}]} = \frac{6.8\times10^{-6}}{8.5\times10^{-4}} = 8.0\times10^{-3}M$

$[CO_3^{2-}] = [Na_2CO_3] = \frac{8.0\times10^{-3}\ mol}{1\ L} \times \frac{106.0\ g}{1\ mol} \times 2.50\ L = \boxed{2.1\ g}$

15. A sample of a saturated solution of MgF_2 was evaporated and the following data table was constructed. Use this data to calculate the value of Ksp for MgF_2 at 25 °C.

Mass of empty evaporating dish	78.5418 g
Mass of evaporating dish and MgF_2 residue after evaporation	78.5434 g
Volume of saturated MgF_2	100.00 mL
Temperature 25.0 °C	

Step 1: Mass of MgF_2 = 78.5434 g − 78.5418 g = 0.0016 g

Step 2: Moles of MgF_2 = $0.0016\ g \times \frac{1\ mol}{62.3\ g} = 2.568\times10^{-5}\ mol$

Step 3: Concentration of MgF_2 = $\frac{2.568\times10^{-5}\ mol}{0.10000\ L} = 2.568\times10^{-4}\ M$

Step 4: Net ionic equation (to determine molar ratios)

$MgF_{2(s)} \rightarrow Mg^{2+} + 2F^-$

$2.568\times10^{-4}M$ $\quad 2.568\times10^{-4}M$ $\quad 5.136\times10^{-4}M$

Step 5: Ksp expression + calculate

$Ksp = [Mg^{2+}][F^-]^2$

$= (2.568\times10^{-4}M)(5.136\times10^{-4})^2$

$Ksp = \boxed{6.8\times10^{-11}}$

16. Calculate the $[Ag^+]$ required to just start precipitation of Ag_2CO_3 in a 0.0030 M solution of $(NH_4)_2CO_3$.

↳ TIP = Ksp

$Ag_2CO_{3(s)} \rightarrow 2Ag^+ + CO_3^{2-}$

TIP = Ksp: $[Ag^+]^2[CO_3^{2-}] = 8.5\times10^{-12}$ ← look up in solubility table

$$[Ag^+] = \sqrt{\frac{Ksp}{[CO_3^{2-}]}} = \sqrt{\frac{8.5\times10^{-12}}{0.0030M}} = \boxed{5.3\times10^{-5}\ M}$$

(2 s.f) 17. A solution is prepared by mixing 20.0 mL of 0.60 M Na_2SO_4 with 60.0 mL of 1.1 M NaOH. Calculate the $[Na^+]$ in the final mixture.

$Na_2SO_4 \rightarrow 2Na^+ + SO_4^{2-}$

~~0.60M~~ ~~1.20M~~ ~~0.60M~~

$$[Na_2SO_4]_{dil} = \frac{C_1V_1}{V_2} = \frac{(0.60)(0.02L)}{(0.08L)} = 0.15M$$

$Na_2SO_4 \rightarrow 2Na^+ + SO_4^{2-}$

0.15M 0.30M 0.15M

$NaOH \rightarrow Na^+ + OH^-$

~~1.1M~~ ~~1.1M~~ ~~1.1M~~

$$[NaOH]_{dil} = \frac{C_1V_1}{V_2} = \frac{(1.1M)(0.06L)}{(0.08L)} = 0.825M$$

$NaOH \rightarrow Na^+ + OH^-$

0.825M 0.825M 0.825M

Final $[Na^+] = 0.30M + 0.825M = \boxed{1.1M}$

(4 s.f) 18. The molar solubility of nickel (II) sulphide is 3.317×10^{-11} M. Calculate the value of the solubility product for nickel (II) sulphide. Show all of your work clearly.

Ksp ←

$NiS_{(s)} \rightarrow Ni^{2+} + S^{2-}$

3.317×10^{-11} 3.317×10^{-11} 3.317×10^{-11} M (1:1 ratio)

$$Ksp = [Ni^{2+}][S^{2-}]$$
$$= (3.317\times10^{-11})(3.317\times10^{-11}M)$$
$$Ksp = \boxed{1.100\times10^{-21}}$$

(3 s.f) 19. A solution of potassium chloride is titrated with 0.200 M silver nitrate solution. The following data table was obtained.

	Trial 1	Trial 2	Trial 3
Initial $AgNO_3$ burette reading (mL)	0.00	5.26	14.63
Final $AgNO_3$ burette reading (mL)	5.26	12.19	19.87
Volume of KCl titrated	25.0	25.0	25.0

Use the information in the data table to calculate the [Cl⁻] in the KCl solution. Show all of your work clearly.

KCl – sample ("unknown") solution → Purpose: to determine $[Cl^-]$ ions in KCl

$AgNO_3$ – titrant or standard ("known") solution

Step 1: Determine volume $AgNO_3$ used for each trial (Final $AgNO_3$ – Initial $AgNO_3$)

	TRIAL 1	TRIAL 2	TRIAL 3
Vol $AgNO_3$ (mL)	5.26	6.93	5.24

Step 2: Average volume: $\frac{5.26 + 6.93 + 5.24}{3} = 5.81\text{ mL } AgNO_3$ or 0.00581 L

Step 3: Calculate moles of $AgNO_3$

moles $AgNO_3 = (C)(V) = (0.200\text{M})(0.00581\text{L}) = 0.001162\text{ mol } AgNO_3$

Step 4: Determine moles of KCl / $[Cl^-]$ using molar ratios

$$AgNO_3 + KCl \rightarrow AgCl_{(s)} + KNO_3$$

$$AgCl_{(s)} \rightarrow Ag^+ + Cl^-$$

mol $AgNO_3$ = mol Ag^+ = mol Cl^- = mol KCl (all 1:1 ratio)

mol Cl^- = 0.001162 mol

Step 5: determine concentration: $[Cl^-] = \frac{mol}{L} = \frac{0.001162\text{ mol}}{0.0250\text{ L}} = 0.4648 = \boxed{0.465\text{ M}}$

← from 25.0 mL of sample solution