1) First, write a balanced net ionic equation:

\[ \text{Ca}^{+2} + \text{SO}_4^{-2} \rightarrow \text{CaSO}_4 \]

We are given a mass of CaSO\(_4\) (1.36 g), and can figure out the moles; these must be the same number of moles of M\(_2\)SO\(_4\), and we can then figure out the M\(_m\) of M\(_2\)SO\(_4\), and the mass and identity of M.

\[
\text{Moles CaSO}_4 = \frac{1.36g}{136.122g/mol} = 9.99 \times 10^{-3} \text{ moles}
\]

Therefore, 1.42 g of M\(_2\)SO\(_4\) must represent 9.99 x 10\(^{-3}\) moles, which gives us a M\(_m\) of \[
\frac{1.42g}{9.99\times10^{-3} \text{ mole}} = 142.13 \text{ g/mol}. \]
This 142.13 g = 2M + 96.062, so 2M = 46.068, and M = \textbf{23.034 g/mole} which means M = Na.

2) So to start with, we know we have Co\(_x\)Cl\(_y\)•nH\(_2\)O, and a bunch of reactions, written as best we can:

\[
\text{Co}_x\text{Cl}_y\cdot n\text{H}_2\text{O} + y\text{AgNO}_3 \rightarrow y\text{AgCl} (s) + \text{Co}_x(\text{NO}_3)_y \text{ (all aqueous except the AgCl)}
\]

\[
\begin{array}{c}
\text{M}_m \\
\text{grams} \\
\end{array}
\begin{array}{c}
143.321 \\
0.308 g \\
\end{array}
\]

The AgCl represents all the chloride in the molecule, so the number of moles of chloride we have is:

\[
\text{Moles Cl}^- = \frac{0.308g}{143.321g/mol} = 2.15 \times 10^{-3} \text{ moles}. \text{ This analysis also tells us that the starting compound is } \left( \frac{0.00215\text{mol}}{0.256g} \right) \times 100 = 29.76\% \text{ Cl} \\
\]

The second equation is a bit more interesting, but still can be written as:

\[
\text{Co}_x\text{Cl}_y\cdot n\text{H}_2\text{O} + \text{NaOH (excess)} \rightarrow \text{Co}_x(\text{OH})_y \rightarrow \text{Co}_2\text{O}_3
\]

\[
\begin{array}{c}
\text{M}_m \\
\text{grams} \\
\end{array}
\begin{array}{c}
165.863 \\
0.145 g \\
\end{array}
\]

The cobalt here represents all the cobalt in the starting compound, and is twice as much, mole-wise, as the number of moles of cobalt(III) oxide. Therefore, we can say:
Moles Co$_2$O$_3$ = \( \frac{0.145 \text{g}}{165.863 \text{g/mol}} \) = 8.742 x 10$^{-4}$ moles, which, similar to above, leads to a % Co of: \( \frac{0.0008742 \text{moles} \times 58.933 \text{g/mole} \times 2 \text{moles Co / Co(III) oxide}}{0.416} \times 100 = 24.77\% \text{ Co.} \)

These percents total 54.53\%, and that appears to account for the Co and Cl present; the rest (45.47\%) must be waters of hydration. Due to the Law of Constant Proportions, we can say that, since water is 11.19\% H, and 88.81\% O, we have 5.09\% H and 40.38\% O.

a) Thus our analysis is **24.77\% Co, 29.76\% Cl, 5.09\% H, and 40.38\% O**.

b) Now to the formula. Using our usual assumption (100 g of compound), we get:

24.77 g Co which is 0.4199 moles Co; 29.76 g Cl which is 0.8394 moles Cl; 5.09 g H which is 5.05 moles H; and 40.38 g O which is 2.524 moles O. Continuing in our usual manner, we get **CoCl$_2$•6H$_2$O**.

c) The equations:

i) CoCl$_2$•6H$_2$O + 2AgNO$_3$ → 2AgCl(s) + Co(NO$_3$)$_2$ + 6H$_2$O(l)

ii) CoCl$_2$•6H$_2$O + 2NaOH → Co(OH)$_2$ + 2NaCl + 6H$_2$O(l)

iii) 4Co(OH)$_2$(s) + O$_2$(g) → 2Co$_2$O$_3$(s) + 4H$_2$O(l) (you may have had to look up the reduction of oxygen.)

3) Here is our general (not balanced!) equation: HCl/HBr + NaOH → NaCl/NaBr + H$_2$O, although we can not use this equation to do the problem.

\[
\begin{align*}
\text{HCl/HBr} + \text{NaOH} & \rightarrow \text{NaCl/NaBr} + \text{H}_2\text{O} \\
100.00 \text{ mL} & \quad 47.14 \text{ mL} \\
\text{unknown } M & \quad 0.1235 \text{ M}
\end{align*}
\]

which means we’ve used 5.893 x 10$^{-3}$ moles OH$^-$ (which is the total moles of H$^+$)

Let x = moles HCl and y = moles HBr; therefore, we can say:

\[
x + y = 5.893 \times 10^{-3} \text{ moles}
\]

We also know that AgCl (M$_m$ = 143.321) and AgBr (M$_m$ = 187.777) are precipitated, with a combined mass of 0.9974 g for the silver halides. Each HX (X=Cl, Br) yields one X$^-$ per HX unit so we can also say:

\[
143.321x + 187.777y = 0.9974 \text{ g}
\]

mass of mass of total
AgCl AgBr salt mass

Therefore, we can substitute for either x or y, and solve. I’ll do solving for y in terms of x:
5.893 \times 10^{-3} - x = y \text{ (from the first equation), which can be substituted into the second equation to get:}

143.321x + 187.777(5.893 \times 10^{-3} - x) = 0.9974 \text{ g}

which simplifies to: 1.1066 - 44.456x = 0.9974. This equation can be restated as:

\[-44.456x = -0.1092, \text{ and } x = \text{ moles HCl} = 2.46 \times 10^{-3}; \text{ therefore, } y = \text{ moles HBr} = 3.44 \times 10^{-3}.\]

However, we were asked for molarities, so dividing by volume ( = 100.00 mL = 0.10000 L) gives molarity of HCl = 0.0246, and molarity of HBr = 0.0344.

(There are, of course, other ways to approach this problem; one involves finding an average atomic mass for the precipitate, then getting percentages of AgCl and AgBr in that mass, and using those data to solve for HCl and HBr; that method is a bit longer, and perhaps tougher, as there are more avenues for error.)

4) You know the procedure; here are just the answers:
   a) 2 \text{ Fe(s)} + 8 \text{ HCl(aq)} \rightarrow 2 \text{ HFeCl}_4(aq) + 3 \text{ H}_2(g)
   b) \text{ Cr(CNS)}_6^{4-}(aq) + 97 \text{ Ce}^{4+}(aq) + 54 \text{ H}_2\text{O(l)} \rightarrow \text{ Cr}^{3+}(aq) + 97 \text{ Ce}^{3+}(aq) + 6 \text{ NO}_3^{-}(aq) + 6 \text{ CO}_2(g) + 6 \text{ SO}_4^{2-}(aq) + 108 \text{ H}^+(aq)
   (This one was quite the toughie!)
   c) 2 \text{ Fe(OH)}_2(s) + \text{ H}_2\text{O}_2(aq) \rightarrow 2 \text{ Fe(OH)}_3(s) \text{ (Sort of reward for the above!)}
   d) 3 \text{ IO}_3^{-}(aq) + 24 \text{ I}^- (aq) + 18 \text{ H}^+(aq) \rightarrow 9 \text{ I}_3^{-}(aq) + 9 \text{ H}_2\text{O(l)}

5) No, I would not expect a precipitation reaction, since the ionic compound that is not an electrolyte is also probably not soluble. Since reactions occur in solution, without solubility, no reaction should be expected.