1. a) Reagent A: ammonia solution complexes formed: middle tube: \([\text{Cu(H}_2\text{O)}_4]\text{(OH)}_2]^{2+}\) last tube: \([\text{Cu(NH}_3)_4(H}_2\text{O)}_2]\) b) Reagent B: sodium hydroxide solution complex formed: \([\text{Cu(H}_2\text{O)}_4]\text{(OH)}_2]^{2+}\) Reagent C: sodium carbonate solution compound formed: \(\text{CuCO}_3\) (These answers could, of course, equally well be the other way around, so that B could be the sodium carbonate solution. Also the reactions would be identical if you used potassium rather than sodium compounds.) c) Reagent D: potassium iodide solution compound formed: \(\text{CuI}\) (There is CuI in both of the tubes containing product. It is produced as a precipitate in a solution of iodine, and takes some time to settle. Sodium iodide would produce the same result.) d) Reagent E: concentrated hydrochloric acid complex formed: \([\text{CuCl}_4]^{-2}\) (There is still unchanged \([\text{Cu(H}_2\text{O)}_4]\text{(OH)}_2]\) present in the tube because the reaction is reversible.)

2. a) This is a redox reaction. The copper has been reduced from a +2 oxidation state to +1. The iodine has been oxidised from a -1 oxidation state to the zero oxidation state in the element.

(Trying to explain this in terms of electron transfer is a bit more of a bother because you would have to construct electron-half-equations. I suppose you might also describe this as a precipitation reaction because a precipitate is formed. That misses the point! The important thing here isn't the formation of the precipitate, but the change of oxidation state.)

b) (i) \[
\text{moles of } \text{Na}_2\text{S}_2\text{O}_3 = \frac{22.5}{1000} \times 0.100
= 0.00225
\]

(You could equally well record this as 2.25 x 10^{-3}, because that’s the way it will appear on your calculator if you used one.)

(ii) \[
\text{2S}_2\text{O}_3^{2-} \text{ (aq)} + \text{I}^{-2} \text{ (aq)} \rightarrow \text{S}_4\text{O}_6^{2-} \text{ (aq)} + 2\text{I}^{-} \text{ (aq)}
\]

2 moles of thiosulphate ions (from 2 moles of sodium thiosulphate) react with 1 mole of iodide. That 1 mole of iodide came from 2 moles of copper(II) ions reacting.

\[
2\text{Cu}^{2+} \text{ (aq)} + 4\text{I}^{-} \text{ (aq)} \rightarrow 2\text{CuI} \text{(s)} + \text{I}^{-2} \text{ (aq)}
\]

So overall, the number of moles of copper(II) ions is the same as the number of moles of sodium
thiosulphate that you used: 0.00225 moles

(iii) There were 0.00225 moles of copper(II) ions in 25 cm$^3$ of the solution you started with – so work out the concentration of the copper(II) ions in mol dm$^{-3}$. The concentration of the copper(II) sulphate solution is, of course, the same as the concentration of copper(II) ions.

\[
\text{concentration of CuSO}_4 = \frac{0.00225 \times 1000}{25} = 0.0900 \text{ mol dm}^{-3}
\]

3. a) The difference is that copper(I) iodide and copper(I) chloride are both insoluble in water, and so produced as solids. Copper(I) sulphate would be soluble in water, and so would produce a solution containing copper(I) ions and sulphate ions. But copper(I) ions undergo disproportionation in solution to give copper(II) ions and a precipitate of copper.

That means that any attempt to produce copper(I) sulphate ends up producing copper(II) sulphate and copper.

b) (i) [CuCl$_2$]

(ii) Add water which removes the extra chloride ion and leaves CuCl as a precipitate. The precipitate has to be separated and dried quickly to prevent it disproportionating.