1. Work out the oxidation state of the named elements:
   a) chlorine in HCl, HClO, NaClO₂, KClO₃, Cl₂O₇, ClO₂
   b) phosphorus in PH₃, PCl₅, H₃PO₄, P₄O₁₀, HPO₃⁻
   c) chromium in Cr, Cr(H₂O)₆³⁺, Na₂CrO₄, Cr₂O₇²⁻

2. In the following equations, state whether the element in **bold** type on the left-hand side has been oxidised or reduced or neither.
   a) \(3\text{Cu} + 8\text{HNO}_3 \rightarrow 3\text{Cu(NO}_3)_2 + 2\text{NO} + 4\text{H}_2\text{O}\)
   b) \(2\text{KBr} + \text{Cl}_2 \rightarrow 2\text{KCl} + \text{Br}_2\)
   c) \([\text{Cu(H}_2\text{O)}_6]^{2+} + 4\text{NH}_3 \rightarrow [\text{Cu(NH}_3)_4(\text{H}_2\text{O})_2]^{2+} + 4\text{H}_2\text{O}\)

3. Work out the equation for the reaction between iron(II) ions and dichromate(VI) ions in acid solution using the following steps as a guide.
   a) Work out the reacting proportions by using the oxidation state changes for iron and chromium using the information:
      
      Iron(II) ions are oxidised to iron(III) ions. Dichromate(VI) ions, Cr₂O₇²⁻, are reduced to chromium(III) ions.
   b) Derive a fully balanced equation by making reasonable assumptions about anything else that might be involved.