OXIDATION STATES

1. Work out the oxidation state of the named elements:

   a) chlorine in HCl, $HClO$, $NaClO_2$, $KClO_3$, Cl_2O_7, ClO_2

   b) phosphorus in PH_3, PCl_5, H_3PO_4, P_4O_{10}, HPO_3^{2-}

   c) chromium in Cr, $Cr(H_2O)_6^{3+}$, Na_2CrO_4, $Cr_2O_7^{2-}$

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) $3\mathbf{Cu} + 8HNO_3 \longrightarrow 3Cu(NO_3)_2 + 2NO + 4H_2O$

   b) $2KBr + \mathbf{Cl_2} \longrightarrow 2KCl + Br_2$

   c) $[\mathbf{Cu}(H_2O)_6]^{2+} + 4NH_3 \longrightarrow [Cu(NH_3)_4(H_2O)_2]^{2+} + 4H_2O$

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, $Cr_2O_7^{2-}$, are reduced to chromium(III) ions.

   b) Derive a fully balanced equation by making reasonable assumptions about anything else that might be involved.