HL Answers to First row d-block elements questions

1. A transition element contains an incomplete d sub-level in one or more of its oxidation states. Zinc loses the two 4s electrons to form zinc(II) compounds. Zn^{2+} has a full 3d level so compounds of zinc(II) do not contain an incomplete d sub-level.

2. All transition elements except chromium and copper contain two 4s electrons. They lose these to form the 2^+ ion. Chromium and copper can also lose one of their 3d electrons (in addition to the 4s electron) to form a 2^+ ion. Copper has the configuration $[Ar]4s^13d^{10}$. Because the full 3d level gives some stability copper can form the Cu^+ ion by just losing its one 4s electron.

3. From the Lewis structure of carbon monoxide it can be seen that both carbon and oxygen possess a non-bonding pair of electrons. The non-bonding pair of electrons on the carbon atom enables it to function as a ligand by forming a coordinate bond with the transition metal atom or ion. Methane does not contain a non-bonding pair of electrons so it cannot function as a ligand (Lewis base).

4. i. Aluminium ($1s^22s^22p^63s^23p^1$) only forms the 3^+ ion to attain a noble gas configuration ([Ne]) as it takes much too much energy to remove a fourth electron from the *n*=2 level. For vanadium there is no large jump in values until the sixth electron is removed. This is because V has the configuration $[Ar]4s^23d^3$ so it can lose its first five electrons relatively easy. Vanadium thus shows oxidation states of +2, +3, +4 and +5.

ii. Superficially it looks as if the same argument should hold true for the first seven electrons of manganese ($[Ar]4s^23d^5$ before a much larger amount of energy is required to remove the eighth electron so manganese can have oxidation states up to +7. However in the case of +5 for vanadium and +6 and +7 for manganese it is not that simple as the V^5 and Mn^{6+} and Mn^{7+} ions are not formed. Compounds such as VO_3^- where vanadium has an oxidation state of +5 and MnO_4^- where manganese has an oxidation state of +7 are complex ions and are not formed by the metal losing five and seven electrons respectively. In fact, very few transition metals form common ions that are greater than M^{3+}. The statement in the syllabus should be viewed with some suspicion.

5. Copper(I) has the electron configuration $[Ar]3d^{10}$ and scandium(III) has the electron configuration [Ar] so neither has any unpaired electrons. For compounds to be paramagnetic they must possess one or more unpaired electrons. In complex ions of most other transition metals the five d orbitals are split and are not completely filled so there is a high probability that at least one of the d electrons will be unpaired.

6. i. $[Fe(CN)_6]^{3-}$ **ii.** $[CuCl_4]^{2-}$ **iii.** $[Co(H_2NCH_2CH_2NH_2)_3]^{3+}$ **iv.** $[Fe(SCN)(H_2O)_5]^{2+}$