Answers to exam-style questions

Topic 3

1 B

2 A

3 C

4 B

5 B

6 C

7 B

8 C

9 B

10 C

11 a i Aluminium and sodium are in the same period in the periodic table, and therefore have the same number of shells of electrons. The shielding from inner shells of electrons is approximately the same in both, but aluminium has a higher nuclear charge (more protons) than sodium, and therefore the outer electrons are attracted more strongly. **[2]**

ii An Al atom forms a 3+ ion by losing its three outer shell electrons. If the Al atom and the Al^{3+} ion are compared, they both have the same number of protons in the nucleus, but the Al^{3+} ion has one fewer shell of electrons. The Al^{3+} ion is therefore smaller than the Al atom. A Cl atom forms a 1− ion by gaining one electron. If the Cl atom and the Cl^- ion are compared, they have the same number of protons in the nucleus, but Cl^- has one extra electron, so there is greater electron–electron repulsion for the same nuclear charge pulling in the electrons. The electron cloud expands and the ion is larger than the atom. **[4]**

iii K forms the K^+ ion, but Cl forms the Cl^- ion. These have the same number of electrons, but K^+ has a higher nuclear charge (19+) than Cl^- (17+), so the electrons are pulled in more strongly and the K^+ ion is smaller. **[2]**

b i First electron affinity: $Cl(g) + e^- \rightarrow Cl^-(g)$
First ionisation energy: $Mg(g) \rightarrow Mg^+(g) + e^-$ **[2]**

ii A chlorine atom is smaller than a bromine atom; so when an electron is accepted into the outer shell of Cl it is closer to the nucleus – and so more strongly attracted to it. **[2]**

iii A chlorine atom is smaller and has a higher nuclear charge than a magnesium atom. However, Cl and Mg have the same number of electron shells, and so approximately the same amount of shielding from inner shells. So, the outer electron is held more tightly in Cl and is more difficult to remove. **[2]**

12 a i $1s^22s^22p^63s^23p^64s^13d^{10}$

ii $1s^22s^22p^63s^23p^63d^9$ **[2]**

b Ligands are negative ions or neutral molecules that possess at least one lone pair of electrons; a lone pair is used to form a coordinate (dative) covalent bond between the ligand and the transition metal ion to form a complex ion. **[2]**

c Cl^- is a negatively charged ligand and four Cl^- ions have a total charge of 4−; the overall charge on the complex ion is 2−; therefore the oxidation number of Cu must be +2. **[1]**

d CuBr contains the Cu^+ ion, which has the electron configuration $1s^22s^22p^63s^23p^63d^{10}$; it has no unpaired electrons and will therefore be diamagnetic. Br^- also has no unpaired electrons. **[2]**

e i The $[Cu(H_2O)_6]^{2+}$ complex ion contains the Cu^{2+} ion, which has a partially filled 3d subshell. The d orbitals are split into two groups in the complex ion. Energy in the form of a certain frequency of visible light is absorbed to promote an electron from the lower set of d orbitals to the upper set. The light transmitted has the complementary colour to the light absorbed. **[3]**

ii $[Cu(NH_3)_2]^+$ contains the Cu^+ ion, which has a full 3d subshell. There is no space available in the upper set of d orbitals to promote an electron to, so light in the visible region of the spectrum cannot be absorbed. **[2]**

iii According to the spectrochemical series, Br^- is a weaker field ligand than H_2O and therefore causes less splitting of the 3d orbitals. The wavelength of light absorbed by $[Cu(H_2O)_5Br]^+$ will therefore be longer than that absorbed by $[Cu(H_2O)_6]^{2+}$; longer wavelength corresponds to lower energy. **[2]**