1. What explains the fact that at room temperature $F_2$ is a gas while $I_2$ is a solid?

The bond in $F_2$ is non-polar so it just has London forces, whereas $I_2$ is polar and therefore has stronger London forces.

2. Write the number of valence electrons possessed by each of the following species:

   a) Ge 4  
   b) Y 3  
   c) W$^{4+}$ 2  
   d) Sn$^{2+}$ 2  
   e) Bi 5  

   Maximum of 2  

3. Which substance has a lower melting temperature: CaBr$_2$ (s) or SrI$_2$ (s)? Explain why.

SrI$_2$ because the ionic bonds are weaker.

4. Write the valence for each: 
   a) S 2  
   b) B 3
   c) Ca 2  
   d) Xe 0  
   e) Ga 3  
   f) Bi 3

5. Write the number of open and closed shells for each of the species given below.

   a) At$^{3+}$ 1, 5  
   b) Ba 1, 5  
   c) I$^{5+}$ 1, 4  
   d) V$^+$ 1, 3

6. Although oxygen and fluorine molecules are roughly the same size, each oxygen atom shares four electrons with its neighbour while each fluorine atom shares only two electrons with its neighbour. Which bond should be stronger? Explain.

The O$_2$ bond is stronger because the shared electrons are distributed more evenly.

7. Which types of bonds or forces increase in strength going down a group in the periodic table? (I only want you to consider bonds/forces between atoms or molecules of the same element)

   London forces

8. Which types of bonds or forces decrease in strength going down a family in the periodic table? (Again, only in terms of bonds/forces between atoms or molecules of the same element)

   Covalent and metallic bonds

9. Provide brief explanations for the following facts:

   a) the melting temperatures of the noble gases increase going down the periodic table.

      Stronger London forces, stronger intermolecular force to overcome.

   b) the reactivity of the alkali metals increases going down the periodic table.

      Valence electron and less nuclear pull, easier to give 'em up.

   c) the reactivity of the noble gases increases going down the periodic table.

      Valence electron is under less and less nuclear pull, easier to get to share.

   d) the melting temperature of the alkali metals decreases going down the periodic table.

      Metallic bonds weaken as atom size increases.

   e) the reactivity of the halogens decreases going down the periodic table.

      As with all non-metals, reactivity corresponds to electronegativity.
10. Define “allotrope”. Name the allotropes of carbon. (Bonus marks possible if they are well drawn and explained!)

11. If a noble gas could form a +1 cation, which noble gas would do so most easily? Why?

Radon. Of all the noble gases, it has the weakest hold on its valence e⁻s.

12. Molecule “M” has closed outer shells as a result of pure covalent bonding within its structure. What type of bond would attract molecules of “M” to each other?

Since it seems the molecule is non-polar, they must attract by London forces.

13. Briefly explain the difference between ionic bonding, covalent bonding and metallic bonding.

Ionic: e⁻ transfer (no sharing)
Covalent: e⁻ sharing between two atoms (1:1)
Metallic: e⁻ shared with all neighboring atoms (not 1:1).

14. Use your periodic table (or the one in Hebden, p 200), to classify each of the bonds as covalent, polar covalent, or ionic. Then predict the formula of the compound based on the valence each type of atom.

a) Ba and Br

Δχ = 1.8: Ionic

Ba Br₂

b) Mg and P

Δχ = 0.9: Polar covalent

Mg₂P₂

c) Si and N

Δχ = 1.2: Polar covalent

Si₃N₄

d) As and F

Δχ = 1.9: Ionic

AsF₃

e) Ga and Se

Δχ = 0.8: Polar covalent

Ga₂Se₃

f) B and Te

Δχ = 0.1: Covalent

B₂Te₃

15. For each pair of compounds, state which should have the higher boiling points. Why?

a) BCl₃ vs. BCl₃F

BCl₃ by dipole-dipole forces.

The London forces in BCl₃ are much weaker.

b) NH₃ vs. SbH₃

NH₃ due to hydrogen bonding.

c) CH₃CH₂CH₃ vs. CH₃CH₂CH₂CH₂CH₃

Hexane has more e⁻-s.

Greater London forces

d) CH₃CH₂CH₃ vs. CH₃OH

Methanol due to hydrogen bonding.