**\*CHEM113: “Cheat sheet” for sequential ordering in explanations – structure and bonding**

The cards of pages 3-9 are colour coded. Print each set, cut out the cards and shuffle them. Then try to lay them out in a logical order that you can use in your explanations. The first 2 pages give a recommended layout, but are not necessarily the only way. Once you can quickly and easily lay the cards in a logical order, practice answering questions using your layout as a guide so that you include enough detail in your answer.

**Periodic trends (see cards for practising the order of explanation)**

1. What are the electron configurations of the atoms?
2. Are valence electrons in the same energy level or a different energy level?
3. Is the distance between the nucleus and the valence/bonding electrons increasing?
4. What is shielding? Is shielding increasing or staying the same?
5. Is the number of protons/nuclear charge increasing?
6. Is the number of protons causing an increase in the attraction between the nucleus and the valence electrons?
7. Is the increased shielding and distance between the valence electrons and the nucleus decreasing the attraction between the nucleus and the valence electrons?
8. Is the radius / ionisation energy / electronegativity increasing?

e.g. Rank O, S and Cl in order of increasing electronegativity and justify your ranking.

**Lewis structures**

1. Count number of valence electrons (remember that Group 18 elements have 8 valence electrons)
2. For cations: Subtract the same number of electrons as in the charge. For anions: Add the same number of electrons as in the charge.
3. Decide which is the central atom (highest valency / least number of atoms of that element) and arrange the other atoms around the central atom
4. Add single bonds by placing 2 electrons between each atom and central atom (you can draw a single line to represent 2 electrons)
5. Arrange remaining valence electrons around outer atoms so each atom has an octet (exception = H, B and Be).
6. If too few electrons for an octet, move as many non-bonding electron pairs to become bonds until each atom has an octet (exception = H, B and Be).
7. If too many electrons for an octet, then add extra electron pairs around central atom. (If central atom is in third period/row of the Periodic Table or higher, they can have more than octet).
8. For ions, include square brackets and the charge

e.g. Draw Lewis structures of ICl4– and HNO3

**Shape and bond angles**

1. What is the central atom?
2. What is the Lewis structure for the molecule/ion?
3. How many regions of negative charge/electron density are around the central atom?
4. What is the electron geometry when these regions of negative charge repel each other?
5. What are the bond angle(s) for this electron geometry?
6. How many bonding regions of electrons are around the central atom?

7. What is the shape of these bonding electron regions?

8. What is the shape of the molecule/ion?

e.g. Explain why ICl4– and ICl3 have similar bond angles.

**\* Originally prepared by Delene Holm as a resource for students at Te Kura and adapted for CHEM113 at VUW**

**Polarity of molecules**

1. Do the bonded atoms have different electronegativities?
2. Are the bonds polar?
3. Do the bonds all have the same polarity?
4. What is the shape of the molecule?
5. Is the shape symmetrical or asymmetrical?
6. Do bond dipoles cancel? / Is there an uneven distribution of electron density? / do polar effects cancel out?
7. Is there a molecular dipole
8. Is the molecule polar or non-polar?

e.g. Explain whether ICl4– and ICl3 are polar or non-polar.

**Comparing Melting points (Δfus*H*) or Boiling points (Δvap*H*)**

1. Are the molecules polar or non-polar?
2. Do any of the molecules have an H atom bonded to N, O or F?
3. What are the all the intermolecular forces for each molecule?
4. How do the molar masses and number of electrons in the molecules compare? (affects temporary dipole forces of attraction)
5. How do the shapes of the molecules compare? (affects temporary dipole forces of attraction)
6. How do the molecular dipoles of the polar molecules compare? (affects permanent dipole forces of attraction)
7. How do the total intermolecular forces for each of the molecules compare?
8. How do the melting/boiling points of each compound compare?

e.g. Account for the relative sizes of the following heats of vaporisation, Δvap*H*: CH3CH2NH2, 29 kJ mol–1;   
HBr, 19 kJ mol–1; Br2, 31.3 kJ mol–1. *M*(CH3CH2NH2) = 45.1 g mol-1, *M*(HBr) = 80.9 g mol-1, *M*(Br2) = 159.8 g mol-1

**Comparing solubility**

1. Are the molecules (solute and solvent) polar or non-polar?
2. Do any of the molecules (solute and solvent) have an H atom bonded to N, O or F?
3. What are the predominate intermolecular forces between the solvent molecules and how much energy (a lot or a little) is needed to break these forces?
4. What are the predominate intermolecular forces between the solute molecules and how much energy (a lot or a little) is needed to break these forces?
5. What sort of intermolecular attractions occur between the solute and solvent molecules?
6. How much energy( a lot or a little) is released by the intermolecular attractions between solvent molecule and solute molecule?
7. Is the energy released by the attractions formed similar to, or greater than the energy required to break the intermolecular attractions?
8. Are the compounds likely to dissolve?

e.g. Compare the solubility of CH3CH2NH2, HBr and Br2 in water.

**Comparing spontaneity of reactions**

1. What is the relationship between spontaneity and entropy?
2. Is the reaction exothermic or endothermic/is heat being given out to the surroundings or taken in from the surroundings?
3. Is the entropy of the surroundings increasing or decreasing?
4. Are particles in the system changing state/becoming more or less ordered?
5. Are there more particles present in the reactants or products in the reaction (system)?
6. Is entropy of the system increasing or decreasing?
7. Is the total entropy (system **and** surroundings) increasing or decreasing?
8. Is the reaction spontaneous or not?

e.g. explain in terms of entropy and enthalpy why vaporisation of methanol can be a spontaneous reaction.

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| **Periodic trends**   1. What are the electron configurations of the atoms? 2. Are valence electrons in the same energy level or a different energy level? 3. Is the distance between the nucleus and the valence/bonding electrons increasing? 4. What is shielding? Is shielding increasing or staying the same? 5. Is the number of protons/nuclear charge increasing? 6. Is the number of protons causing an increase in the attraction between the nucleus and the valence electrons? 7. Is the increased shielding and distance between the valence electrons and the nucleus decreasing the attraction between the nucleus and the valence electrons? 8. Is the radius / ionisation energy / electronegativity increasing? | **Periodic trends**  Are the valence electrons in the same energy level or a new energy level? | **Periodic trends**  Is the distance between the nucleus and the valence/bonding electrons increasing? |
| **Periodic trends**  Is the number of protons/nuclear charge increasing? | **Periodic trends**  What is shielding?  Is shielding increasing or staying the same? | **Periodic trends**  What are the electron configurations of the atoms? |
| **Periodic trends**  Is the number of protons causing an increase in the attraction between the nucleus and the valence electrons? | **Periodic trends**  Is the increase in shielding and in the distance of the valence electrons from the nucleus decreasing the attraction between the nucleus and  valence electrons? | **Periodic trends**  Is the radius / ionisation energy / electronegativity increasing? |

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| **Periodic trends**  O: 1*s*22*s*22*p*4  S: 1*s*22*s*22*p*63*s*23*p*4  Cl:1*s*22*s*22*p*63*s*23*p*5 | **Periodic trends**  S and Cl have valence electrons in the same shell.  O has valence electrons in a shell closer to the nucleus | **Periodic trends**  The distance between the nucleus and valence electrons is much smaller in O due fewer energy levels, but is very similar in S and Cl. |
| **­­Periodic trends**  Shielding is caused by electrons in the inner (core) energy levels decreasing the attraction between the nucleus and valence electrons. So the shielding effect for Cl and S is the same as they both have 2 core energy levels. O only has 1 core energy level so has less shielding. | **Periodic trends**  O has 8 protons  S has 16 protons and Cl has 17 protons  …so can rank elements as O < S < Cl in order of increasing nuclear charge. | **Periodic trends**  Hence order of increasing electronegativity is S < Cl < O |
| **Periodic trends**  Comparing S and Cl: increased nuclear charge in Cl means a greater attraction for the valence electrons.  Comparing S and O: increased nuclear charge in S does not increase attraction of valence electrons | **Periodic trends**  Comparing S and O: increased shielding and distance of the valence electrons from the nucleus means a smaller attraction of valence electrons of S than O, in spite of increased nuclear charge | **Periodic trends** |

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| **Lewis structures**  **Count number of valence electrons**  (remember that Group 18 elements have 8 valence electrons) | **Lewis structures**  **Decide which is the central atom**  (highest valency / least electronegative element) | **Lewis structures**  **Arrange the other atoms around the central atom** |
| **Lewis structures**  **Add single bonds by placing 2 electrons between each atom and central atom** | **Lewis structures**  **Arrange remaining valence electrons around outer atoms so each atom has an octet**  (exceptions: H, B and Be). | **Lewis structures**  **If too few electrons, for an octet move as many non-bonding electron pairs to form a bond until each atom has an octet**  (exceptions: H, B and Be). |
| **Lewis structures**  **If too many electrons for an octet, add extra electrons around central atom.**  If central atom is in third period/row of the Periodic Table or higher, then can have more than octet. | **Lewis structures**  **For cations: Subtract the same number of electrons as in the charge.**  **For anions: Add the same number of electrons as in the charge.** | **Lewis structures**   1. Count number of valence electrons 2. For cations: Subtract the same number of electrons as the charge. For anions: Add the same number of electrons as the charge. 3. Decide which is the central atom (highest valency / least number of atoms of that element) and arrange the other atoms around the central atom 4. Add single bonds by draw a line between each atom and central atom (represents 2 electrons) 5. Arrange remaining valence electrons around outer atoms so each atom has an octet (exception = H, B and Be). 6. If too few electrons for an octet, move as many non-bonding electron pairs to become bonds until each atom has an octet (exception = H, B and Be). 7. If too many electrons for an octet, then add extra electron pairs around central atom. 8. For ions, include square brackets and the charge |

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| **Lewis structures**  **ICl4– : (5 × 7) + 1 = 36**  **POCl3 : 5 + 6 + 3(7) = 32** | **Lewis structures**  **ICl4– : I**  (least number of atoms of that element)  **POCl3 : P**  (least electronegative) | **Lewis structures** |
| **Lewis structures** | **Lewis structures** | **Lewis structures**  **ICl4– : 4 electrons still left**  **Add these to the central atom** |
| **Lewis structures**  **POCl3 : 32 electrons used**  **But O does not have an octet so shift an electron pair to make a double bond** | **Lewis structures**  **ICl4- has an extra electron added to account for the negative charge** | **Lewis structures** |

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| **Shape and bond angles**  **What is the central atom?** | **Shape and bond angles**  **What is the Lewis structure for the molecule/ion?** | **Shape and bond angles**  **How many regions of negative charge/electron density are around the central atom?** |
| **Shape and bond angles**  **What is the electron geometry when these regions of negative charge repel each other to minimise repulsion?** | **Shape and bond angles**  **What are the bond angle(s) for this electron geometry?** | **Shape and bond angles**  **How many bonding regions of electrons are there?** |
| **Shape and bond angles**  **What is the shape of these bonding regions?** | **Shape and bond angles**  **What is the shape of the molecule/ion (name and diagram)?** | **Shape and bond angles**   1. What is the central atom? 2. What is the Lewis structure for the molecule/ion? 3. How many regions of negative charge/electron density are around the central atom? 4. What is the electron geometry when these regions of negative charge repel each other? 5. What are the bond angle(s) for this electron geometry? 6. How many bonding regions of electrons are around the central atom?   7. What is the shape of these bonding electron regions?  8. What is the shape of the molecule/ion? |

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| **Shape and bond angles**  **ICl4– : I**  (least number of atoms of that element)  **ICl3 : I**  (least number of atoms of that element) | **Shape and bond angles** | **Shape and bond angles**  **ICl4– has 6 regions of electron density around the central I atom.**  **ICl3 has 5 regions of negative charge around the central I atom.** |
| **Shape and bond angles**  **To minimise repulsion:**  **the 6 regions of electrons in ICl4– have an octahedral geometry and**  **the 5 regions of electrons in ICl3 have a trigonal planar geometry** | **Shape and bond angles**  **The octahedral geometry has bond angles of 90°.**    **The trigonal bipyramidal shape as an angle of 90° between the axial and equatorial planes and bond angles of 120° within the equatorial plane.** | **Shape and bond angles**  **ICl4– has 4 bonding regions of electrons around the central I atom.**  **ICl3 has 3 bonding pairs of electrons around the central I atom.** |
| **Shape and bond angles**  **ICl4– bonding regions take a square planar shape.**  **ICl3 bonding regions take a T-shape (the non-bonding electrons are in the equatorial plane as this gives minimum repulsion since lone pairs occupy more space than bond pairs).** | **Shape and bond angles** | **Shape and bond angles**  **ICl4– is square planar with bond angles of 90o.**    **ICl3 is a T-shape with bond angles close to 90o.** |

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| **Polarity of molecules**  **Do the bonded atoms have different electronegativity?** | **Polarity of molecules**  **Are the bonds polar?** | **Polarity of molecules**  **Do the bonds all have the same polarity?** |
| **Polarity of molecules**  **What is the shape of the molecule?** | **Polarity of molecules**  **Is the shape symmetrical or asymmetrical?** | **Polarity of molecules**  **Do bond dipoles cancel?**  (or Is there an uneven distribution of electron density? / do polar effects cancel out?) |
| **Polarity of molecules**  **Is there a molecular dipole?** | **Polarity of molecules**  **Is the molecule polar or non-polar?** | **Polarity of molecules**   1. Do the bonded atoms have different electronegativities? 2. Are the bonds polar? 3. Do the bonds all have the same polarity? 4. What is the shape of the molecule? 5. Is the shape symmetrical or asymmetrical? 6. Do bond dipoles cancel? / Is there an uneven distribution of electron density? / do polar effects cancel out? 7. Is there a molecular dipole 8. Is the molecule polar or non-polar? |

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| **Polarity of molecules**  **Cl is more electronegative than I as it is higher up Group 17.** | **Polarity of molecules**  **This means that the I–Cl bonds are polar** | **Polarity of molecules**  **All the bonds are the same so have the same polarity** |
| **Polarity of molecules**  **ICl3 is T-shaped** | **Polarity of molecules**  **This shape is asymmetrical/** | **Polarity of molecules**  **The shape of the molecule means that the bond dipoles are arranged in such a way that they do not cancel each other out** |
| **Polarity of molecules**  **The molecule has a dipole** | **Polarity of molecules**  **ICl3 is a polar molecule** | **Polarity of molecules** |

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| **Melting points (Δfus*H*) or Boiling points (Δvap*H*)**   1. Are the molecules polar or non-polar? 2. Do any of the molecules have an H atom bonded to N, O or F? 3. What are the all the intermolecular forces for each molecule? 4. How do the molar masses and number of electrons in the molecules compare? (affects temporary dipole forces of attraction) 5. How do the shapes of the molecules compare? (affects temporary dipole forces of attraction) 6. How do the molecular dipoles of the polar molecules compare? (affects permanent dipole forces of attraction) 7. How do the total intermolecular forces for each of the molecules compare? 8. How do the melting/boiling points of each compound compare? | **Melting points (Δfus*H*) or Boiling points (Δvap*H*)**  **Are the molecules polar or non-polar?** | **Melting points (Δfus*H*) or Boiling points (Δvap*H*)**  **Do any of the molecules have a H atom bonded to N, O or F?** |
| **Melting points (Δfus*H*) or Boiling points (Δvap*H*)**  **What are the all the different types of intermolecular forces that each molecule can form?** | **Melting points (Δfus*H*) or Boiling points (Δvap*H*)**  **How do the molar masses and number of electrons in the molecules compare? (affects temporary dipole forces of attraction)** | **Melting points (Δfus*H*) or Boiling points (Δvap*H*)**  **How do the shapes of the molecules compare? (affects temporary dipole forces of attraction)** |
| **Melting points (Δfus*H*) or Boiling points (Δvap*H*)**  **How do the molecular dipoles of the molecules compare? (affects permanent dipole forces of attraction)** | **Melting points (Δfus*H*) or Boiling points (Δvap*H*)**  **How does the total intermolecular force for each of the molecules compare?** | **Melting points (Δfus*H*) or Boiling points (Δvap*H*)**  **How do the melting/boiling points/Δfus*H*/Δvap*H* of the compounds compare?**  (Note that a compound has a melting point, not a molecule) |

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| **Melting points (Δfus*H*) or Boiling points (Δvap*H*)**  **CH3CH2NH2 and HBr are polar.**  **Br2 is non-polar** | **Melting points (Δfus*H*) or Boiling points (Δvap*H*)**  **CH3CH2NH2 has two H atoms covalently bonded to a very electronegative N atom, so these H atoms can hydrogen bond with N atoms in a different molecule.** | **Melting points (Δfus*H*) or Boiling points (Δvap*H*)**  **Br2 forms only temporary dipole attractions**  **HBr forms temporary & permanent dipole attractions CH3CH2NH2 forms temporary & permanent dipole attractions and H-bonds** |
| **Melting points (Δfus*H*) or Boiling points (Δvap*H*)**  Br2 has the largest molar mass and hence the largest electron cloud and forms the strongest temporary dipole attractions.  HBr has a lower molar mass / less electrons so has weaker temporary dipole attractions and CH3CH2NH2 has the smallest molar mass /smallest electron cloud so the weakest temporary dipole attractions. | **Melting points (Δfus*H*) or Boiling points (Δvap*H*)**  The shape of CH3CH2NH2 is longer than for the other two molecules and this will slightly increase the strength of these temporary dipole attractions | **Melting points (Δfus*H*) or Boiling points (Δvap*H*)**  **As N is more electronegative than Br, the molecular dipole in CH3CH2NH2 is likely to be bigger than that in HBr** |
| **Melting points (Δfus*H*) or Boiling points (Δvap*H*)**  **Since CH3CH2NH2 has the highest Δvap*H*, it will have the strongest intermolecular forces. This will be due to the presence of strong H-bonds between the molecules.**  **Similarly, Br2 has stronger intermolecular forces than HBr in spite of HBr being polar. This will be due to the much larger electron cloud in Br2.** | **Melting points (Δfus*H*) or Boiling points (Δvap*H*)**  **The stronger the total force of attraction the more energy is required to break those forces of attraction.** | **Melting points (Δfus*H*) or Boiling points (Δvap*H*)** |

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| **Solubility**  **Are the molecules (solute and solvent) polar or non-polar?** | **Solubility**  **Do any of the molecules (solute and solvent) have a H atom bonded to N, O or F?** | **Solubility**  **What are the predominate intermolecular forces between the solvent molecules** |
| **Solubility**  **Is the amount of energy needed break these forces large or small?** | **Solubility**    **What are the predominate intermolecular forces between the solute molecules and is the amount of energy needed break these forces large or small?** | **Solubility**  **What sort of intermolecular attractions occur between solute & solvent molecules and is the amount of energy released when these attractions form large or small?**  **?** |
| **Solubility**  **Are the compounds likely to dissolve?** | **Solubility**  **Is the energy released by the attractions formed similar to, or greater than, the energy required to break the intermolecular attractions?** | **Solubility**   1. Are the molecules (solute and solvent) polar or non-polar? 2. Do any of the molecules (solute and solvent) have an H atom bonded to N, O or F? 3. What are the predominate intermolecular forces between the solvent molecules 4. Is the amount of energy needed to break these forces large or small? 5. What are the predominate intermolecular forces between the solute and is the amount of energy needed to break these forces large or small? 6. What sort of intermolecular attractions occur between the solute and solvent molecules and is the amount of energy needed to break these forces large or small? 7. Is the energy released by the attractions formed similar to, or greater than the energy required to break the intermolecular attractions? 8. Are the compounds likely to dissolve? |

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| **Solubility**  **H2O (solvent), and**  **CH3CH2NH2 (solute) are polar**  **Br2(solute) is non-polar.** | **Solubility**  **H2O molecules can H-bond to each other as they have H atoms bonded to very electronegative O atoms. Also CH3CH2NH2 molecules have 2 H atoms bonded to very electronegative N atoms so can H-bond to each other electronegative N atoms.** | **Solubility**  **CH3CH2NH2 has temporary & permanent dipole attractions and H-bonds between molecules.**  **Br2 only has only temporary dipoles.** |
| **Solubility**  **Less energy is needed to separate the Br2 molecules from each other than to separate**  **the CH3CH2NH2 molecules from each other.** | **Solubility**  **In the solvent, H2O there are temporary and permanent dipole attractions and H-bonds between the molecules.**  **A large amount of energy is needed to break these intermolecular forces** | **Solubility**  **CH3CH2NH2 has 2 H atoms covalently bonded to very electronegative N so can form hydrogen bonds with H2O molecules.**  **Br2 will only form dipole – dipole attractions to water molecules.** |
| **Solubility**  **A large amount of energy is released when H-bonds form between H2O and CH3CH2NH2 molecules.**  **Only a small amount of energy is released when dipole – dipole attractions are formed between to H2O and Br2 molecules.** | **Solubility**  **As the energy required to break solvent-solvent and solute-solute attractions is similar to energy released when new solute-solvent attractions form, CH3CH2NH2 and will be soluble in water.**  **Insufficient energy is released by H2O-Br attractions to break the stronger solvent-solvent and solute-solute attractions, so Br2 is insoluble in water.** | **Solubility** |

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| **spontaneity of reactions**  **What is the relationship between spontaneity and entropy?** | **spontaneity of reactions**  **Is the reaction exothermic or endothermic?**  **Is heat being given out to the surroundings or taken in from the surroundings?** | **spontaneity of reactions**  **Is entropy of the surroundings increasing or decreasing?** |
| **spontaneity of reactions**  **Are the states of the reactants different from the products / are**  **are particles in the system becoming more ordered?** | **spontaneity of reactions**  **Are there more particles present in the reactants or products in the reaction (system)?** | **spontaneity of reactions**  **Is entropy of the sytem increasing or decreasing?** |
| **spontaneity of reactions**  **Is the total entropy, Δ*Suniverse*  (Δ*Ssystem* + Δ*Ssurroundings)* increasing or decreasing?** | **spontaneity of reactions**  **Is the reaction spontaneous or not?** | **spontaneity of reactions**   1. What is the relationship between spontaneity and entropy? 2. Is the reaction exothermic or endothermic/is heat being given out to the surroundings or taken in from the surroundings? 3. Is the entropy of the surroundings increasing or decreasing? 4. Are particles in the system changing state/becoming more or less ordered? 5. Are there more particles present in the reactants or products in the reaction (system)? 6. Is entropy of the system increasing or decreasing? 7. Is the total entropy (system **and** surroundings) increasing or decreasing? 8. Is the reaction spontaneous or not? |

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| **spontaneity of reactions**  **When the total entropy of a reaction (system + surroundings) is positive then the reaction will be spontaneous**  **Δ*Suniverse =* Δ*Ssystem* + Δ*Ssurroundings)*** | **spontaneity of reactions**  **Vaporisation of methanol is an endothermic reaction as heat is taken from the surroundings is change CH3OH*(ℓ)* to CH3OH*(g)*** | **spontaneity of reactions**  **The surroundings will cool down so the entropy of the surroundings will decrease** |
| **spontaneity of reactions**  **Methanol is changing from liquid to a gas so the particles become more disordered.** | **spontaneity of reactions**  **The number of particles does not change as the state changes.** | **spontaneity of reactions**  **Entropy in the system is increasing** |
| **spontaneity of reactions**  **The total enthalpy, system + surroundings, is increasing so the positive entropy of the system must be sufficient to overcome the negative entropy of the surroundings** | **spontaneity of reactions**  **The reaction is spontaneous** | **spontaneity of reactions** |