

Things to remember in the last hour before the exam: Level 2 Structure & Energy

(This is not a revision sheet – you’ve done that by now – it’s a list of things you might want to remind yourself about ...)

1. Types of solids.

Ionic – between metal & non-metal (exception AlCl_3 covalent molecular); particles = IONS, attraction = ionic / electrostatic bond

Covalent – between non-metal and non-metal.

- COVALENT NETWORK: particles = atoms; attraction = covalent bond; E.g. diamond, graphite and silicon dioxide SiO_2 .

- MOLECULAR: particles = molecules; attraction = weak intermolecular; E.g. H_2O , I_2 , CO_2 .

Metallic – bonding between metal atoms; particles involved = atoms; attraction = metallic bond; E.g. Ag.

2. Bonds:

- Metallic: attraction between loosely held valence electrons & positively charged nuclei of neighbouring atoms. (Or “metal cations in sea of delocalised electrons” BUT NCEA examiners haven’t liked this definition very much in the past).

- Ionic: electrostatic attraction between oppositely charged ions.

- Covalent: bond in which one or more pairs of electrons are shared by two atoms. It’s intramolecular (between the atoms within the molecule).

- Intermolecular: weak attraction between molecules (intermolecular; inter = between)

3. Explaining electrical conductivity. (1) metals (solid/molten) – delocalised electrons are free to move (2) ionic substances (molten/dissolved in water (aq)) – ions are free to move. (3) graphite – delocalised electrons are free to move.

4. Properties.

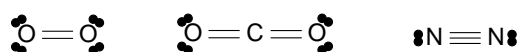
- Hardness – indicates a strong metallic, covalent or ionic bond – much energy needed to overcome.
- Brittle – indicates ionic bond (when like charged ions line up they repel)
- High m.pt., b.pt – indicates strong attraction between atoms or ions (metallic, covalent or ionic bond).
- Low m.pt., b.pt, soft – indicates weak attraction between molecules (molecular covalent)
- Malleable, ductile – indicates the non-directional attraction due to metallic bond.
- Conducts electricity – has “mobile charge carriers” – delocalised electrons in metals (and graphite). Charge carriers are ions in molten or aqueous ionic substances. (In solid, ions NOT free to move).
- Solubility: non-polar substances dissolve in non-polar solvents as similar weak intermolecular forces exist. Polar and ionic solids dissolve in polar solvents (e.g. H_2O) due to attraction between charged particles. (DO NOT call ionic solids “polar” – they are not!)

5. Lewis structures. A pair of electrons are drawn as $\bullet\bullet$ or $\times\times$ or — ; don’t draw $\bullet|\bullet$ OR $\times|x$.

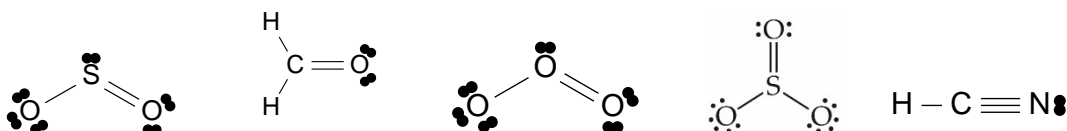
6. The number of valence electrons an atom has = its group number or group number – 10. E.g. Al – group 13. $13 - 10 = 3$ valence electrons. (or work it out 2, 8, 3).

7. BeCl_2 and BCl_3 are electron deficient (don’t have octet around Be or B, the central atom).

8. Some common molecules with multiple bonds are

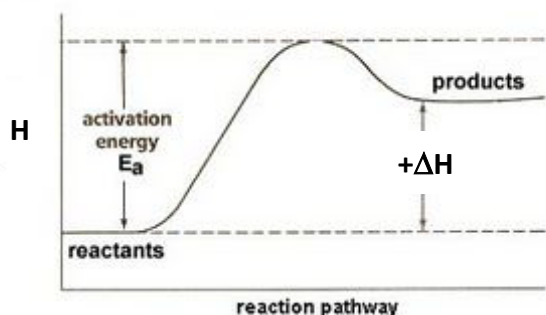


9. A few others you might be asked to draw (with multiple bonds) might be worth memorising!

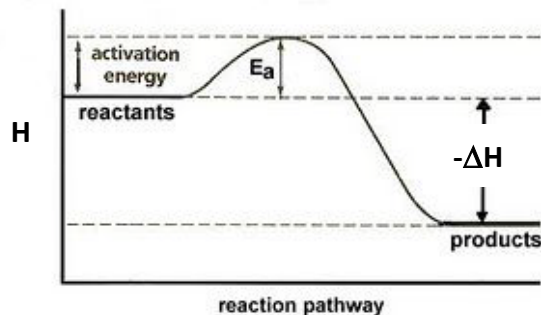


10. Bond angles: Based on repulsion of regions of negative charge. 2 regions – linear 180° . 3 regions – trigonal planar 120° . 4 regions tetrahedral – 109° . Shapes of molecules are based around repulsion of regions of negative charge around central atom, bonding and non-bonding. Shape based on atoms.
- Linear (bond angle 180°)
 - V-shaped (either bond angle 120° or 109°)
 - Trigonal planar (bond angle 120°)
 - Trigonal pyramidal (bond angle 109°)
 - Tetrahedral (bond angle 109°)
11. Electronegativity: Ability of atoms in a bond to attract electrons to themselves. Inc \rightarrow periodic table, inc \uparrow a group. Remember (Most Electronegative)...F O N/Cl S C H ... (less electronegative).
12. Show EN with $\delta+$ and $\delta-$ above atoms OR show dipole \rightarrow with $+$ end over $\delta+$, \rightarrow end over $\delta-$.
13. Polarity of molecules – predicting if molecule is polar or not.
- Polar molecules: contain polar bonds AND lack of molecule symmetry means dipoles do not cancel out. e.g. H_2O
 - Non-polar: contain polar bonds BUT molecule symmetry means dipoles do cancel out e.g. $\text{O}=\text{C}=\text{O}$: OR do not contain polar bonds e.g. if 2 atoms of same electronegativity are bonded. e.g. $\text{Cl}-\text{Cl}$.
14. Energy: H is enthalpy.
- Endothermic – energy absorbed from surroundings. Products have more energy than reactants. $+\Delta H$. Bond breaking is endothermic.
 - Exothermic – energy released to surroundings. Products have less energy than reactants. $-\Delta H$. Bond making is exothermic.
 - E_a = activation energy: Catalysts LOWER E_a but don't change H of reactants/products or ΔH .

ENDOTHERMIC



EXOTHERMIC



15. Change of state: solid \rightarrow liquid \rightarrow gas: Endothermic as bond breaking; Change of state: gas \rightarrow liquid \rightarrow solid: Exothermic as bond making.
16. Bond breaking & making calculations: remember break $+\Delta H$, make $-\Delta H$. Find overall enthalpy change by adding break and make (as long as you have the signs correct the answer will be!); Units are kJ or kJ mol^{-1} . Remember to convert $\text{C}=\text{C}$ to $\text{C}-\text{C}$ break the $\text{C}=\text{C}$ and make a $\text{C}-\text{C}$.
17. Check what question asks! If asked for energy released (or absorbed), just give an answer with units AND NO SIGN, e.g. 300 kJ released (OR 300 kJ absorbed). If Q. asks for enthalpy change then you need a sign e.g. $\Delta H = -300 \text{ kJ}$ (or $\Delta H = +300 \text{ kJ}$).
18. $n=m/M$: n is amount (mol), m is mass (g) and M is molar mass (g mol^{-1}).
19. $\Delta_r H$ can have units kJ or kJ mol^{-1} . The "per mole" refers to a mole of equation! (Don't stress about this!)