Types of solids

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

Covalent – between non-metal & non-metal.

- COVALENT NETWORK; particles = atoms; attraction = covalent bond; E.g. diamond, graphite and silicon dioxide SiO_2.
- COVALENT MOLECULAR; particles = molecules; attraction = weak intermolecular forces; E.g. H_2O, I_2, CO_2.

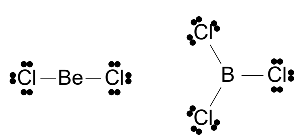
Metallic – bonding between metal atoms; particles involved = atoms; attraction = metallic bond; E.g. Ag

Lewis structures

Show only valence electrons. The number of valence electrons an atom has = its group # or group # – 10. E.g. Al – group 13. 13 – 10 = 3 valence electrons. (Or work it out from Aluminium's atomic number of 13 which gives an electron arrangement of 2, 8, 3).

Pairs of electrons are drawn as ●● or x x

Simple molecules have no more than four electron pairs about any atom (including multiple-bonded species). Be e.g. in $BeCl_2$ and B e.g. in BCl_3 are electron deficient (don't have octet around Be or B, the central atom).



Usually, the central atom is already known. Otherwise the atom with the lowest electronegativity is the central atom.

NOTE: H can never be a central atom.

Defining bond types

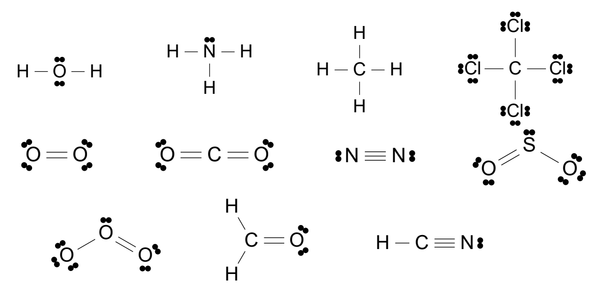
Metallic: attraction between loosely held valence electrons & positively charged nuclei of neighbouring atoms. (Or "metal cations in sea of delocalised electrons").

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 atoms within the molecule).

Intermolecular: weak attraction between molecules (inter = between)

Other useful Lewis diagrams

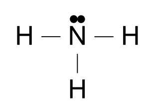


Bond angles and shapes of molecules

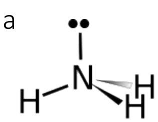
Based on repulsion of regions of electron density. 2 regions – linear 180°, 3 regions – trigonal planar 120°, 4 regions tetrahedral – 109°.

Shapes of molecules are based around repulsion of regions of electron density, bonding and non-bonding BUT ultimately depend on the positions of the atoms.

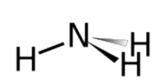
E.g. NH_3



- 4 regions of electron density around the central N atom which repel each other to get as far away from each other as possible, taking up the *arrangement* of a regular tetrahedron.



- 3 regions are bonding, one non-bonding – and so the shape of the molecule is trigonal pyramid / trigonal pyramidal, with bond angle of approximately 109°.



Explaining properties

Hardness – indicates strong metallic, covalent or ionic bond – much energy needed to overcome attraction.

High m.pt, b.pt – indicates strong attraction between atoms or ions (metallic, covalent network or ionic bond).

Brittle – indicates ionic bond (when like force causes like charged ions to line up, they repel)

Low m.pt, b.pt, soft - indicates weak attraction / weak intermolecular forces 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) OR ions in molten or aqueous ionic substances. (NOTE: In solid ionic substances the ions are NOT free to move and so these solids are insulators).

Solubility - non-polar substances e.g. I_2 dissolve in non-polar solvents e.g. hexane, because similar weak intermolecular forces exist between I_2 and hexane molecules as did between I_2 and I_2 molecules and between cyclohexane and cyclohexane molecules.

Polar covalent molecules and many 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!)

Polar or non polar molecules

To predict if a bond is polar, consider electronegativity, the ability of atoms *in a bond* to attract electrons to themselves. Inc → periodic table, inc ↑ a group. Remember (most)…F O N/Cl S/C H (less). Show EN with δ+ and δ- above atoms.

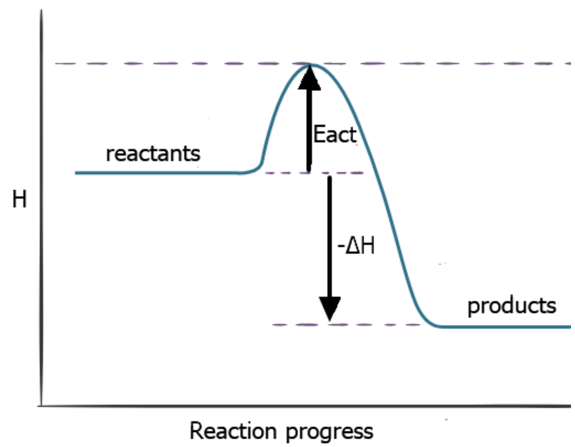
Predicting if molecule is polar or not. *Polar molecules*: contain polar bonds AND their lack of molecule symmetry means dipoles do not cancel out, e.g. H_2O. *Non-polar molecules*: usually contain polar bonds BUT the molecule symmetry means dipoles do cancel out e.g. O=C=O ; OR do not contain polar bonds e.g. where atoms of same electronegativity are bonded, e.g. Cl-Cl.

Exothermic and endothermic reactions

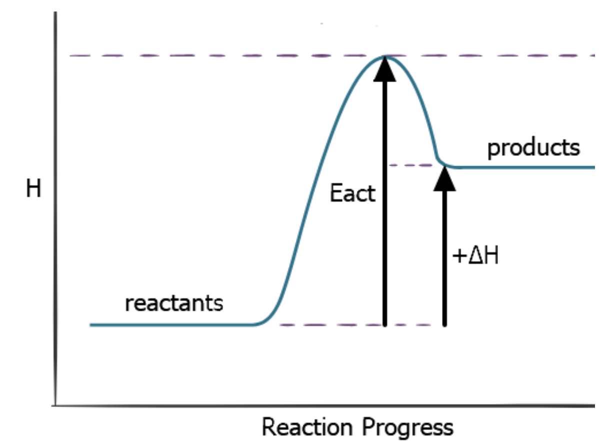
In an exothermic reaction, the reactants are at a higher energy level as compared to the products. The products are more stable than the reactants. Overall Δ_rH for the reaction is negative; energy is released in the form of heat.

In the case of an endothermic reaction, the reactants are at a lower energy level compared to the products. The products are less stable than the reactants. The overall Δ_rH for the reaction is positive, i.e., energy is absorbed from the surroundings.

Exothermic



Endothermic



Enthalpy changes associated with the making and breaking of chemical bonds

Average bond enthalpies have the units of kJ mol^{-1}. They show the energy required to break 1 mol of a particular bond.

Average bond enthalpies are always listed as positive (+) numbers.

Δ_rH^o = ∑(bonds broken) + ∑(bonds formed) where ∑ means "the sum of"

You will need to decide if the value is + or - depending upon whether bonds are being broken (+) or formed (-).

To break bonds is endothermic; to make the same bonds is exothermic

e.g if O=O → O + O +498 kJ mol^{-1}, then O + O → O=O -498 kJ mol^{-1}

Note: To turn C=C into C-C, you must break C=C and then make C-C.

Exothermic and endothermic reactions including energy (enthalpy) changes associated with differing amounts of substances

Calculations of energy changes using Δ_rH and reaction stoichiometry **n= cV**

e.g. $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$ $\Delta H = –890$ $kJmol^{-1}$. When 1 mol of $CH_4(g)$ is completely burnt $\Delta_rH = –890$ $kJmol^{-1}$ OR 890 kJ of energy are released. $M(CH_4) = 16.0$ g mol^{-1}.

When 16.0 g of $CH_4(g)$ is completely burnt $\Delta_rH = –890$ $kJmol^{-1}$ OR 890 kJ of energy are released. (Don't say -890 kJ of energy are released).

What is the enthalpy change when 46.2 g of methane is burnt? Either do

- by ratio: 16.0 g of $CH_4(g)$ $\Delta_rH = –890$ $kJmol^{-1}$, so 46.2 g = 46.2/16.0 x -890 = -2569 $\Delta_rH = –2570$ kJ (3 s.f.)
- by mol: n=m/M n = 46.2/16.0 = 2.89 mol. For 1 mol, $\Delta_rH = –890$ $kJmol^{-1.}$ For 2.89, Δ_rH = 2.89 x -890 = 2570 kJ (3 s.f.)

Changes of state

Solid → Liquid → Gas

Endothermic; bond breaking.

Gas → Liquid → Solid

Exothermic; bond making