Trends- *Across a period*, the valence electrons / bonding electrons are found in the same energy level with the same e^- - e^- repulsion (shielding) from inner energy levels.

The # of protons increases across a period / nuclear charge increases. This means the electrostatic attraction between the positive nucleus and the valence electrons / bonding electrons increases across a period, and therefore:

- the atomic radius decreases.
- more energy is required to remove the outermost valence electrons, so first IE increases.
- bonding electrons are more strongly attracted to the nucleus, so electronegativity increases.

Down a group…. valence electron(s) in an energy level further from the nucleus with greater repulsion/shielding from inner energy levels. Although # of protons increases down a group, effect is offset by the increasing distance between nucleus and valence electrons. So, electrostatic attraction between the positive nucleus and valence electrons decreases, and therefore: atomic radius increases while 1st ionisation energy & electronegativity decrease down a group.

Electronegativity – ability of an atom in a bond to attract the (bonding) electrons towards itself.

1st Ionisation energy – the energy to remove one mol of electrons from one mol of atoms in the gaseous state.

$X(g) \rightarrow X^+(g) + e^-$ *(where x is symbol for any element)*

Entropy – the degree of dispersal of energy/matter

Spontaneous means reaction/process just happens (or carries on once it starts) e.g. salt dissolves, alcohol burns.

To be spontaneous ΔS_{total} must be positive

$$\Delta S_{total} = \Delta S_{system} + \Delta S_{surroundings}$$

ΔS_{system} is the 'reaction' / thing with a → Consider the states AND the number of particles

$\Delta S_{surroundings}$ is the surroundings! Need to consider the enthalpy change ΔH of the system (is it exo or endo?) to see if surroundings heat up (ΔS_{sur}↑) / cool down (ΔS_{sur}↓).

Weak intermolecular attractions

TD-TD: consider size of molecule = larger molecule = more electrons = bigger electron cloud = more polarisable = more/stronger TD-TD attractions. Also consider shape of molecule (linear = more/ branched = less SA for attractions to operate over).

PD-PD: polar molecules, molecules with overall dipole.

Hydrogen bonding: Between molecules with H bonded to F, O or N e.g. NH_3 ✓ but CH_3COCH_3 ✗

Relate the weak intermolecular attractions to mpt, bpt, $\Delta_{fus}H$, $\Delta_{vap}H$. Higher values = stronger attractions between molecules = more energy needed to be put in to overcome them!

Size of atoms vs ions

Metal atoms < metal ions as lose whole energy level. Nonmetal ions > nonmetal atoms as extra electron(s) in valence shell inc. the e^- - e^- repulsion.

Shape/polarity

Around the central atom there are [#] regions of electron density. Repulsion between these regions of electron density results in a [name] electron geometry to maximise separation/minimise repulsion. Since there are [#] bonding regions and [#] non-bonding regions the overall molecule/ion shape is [name].

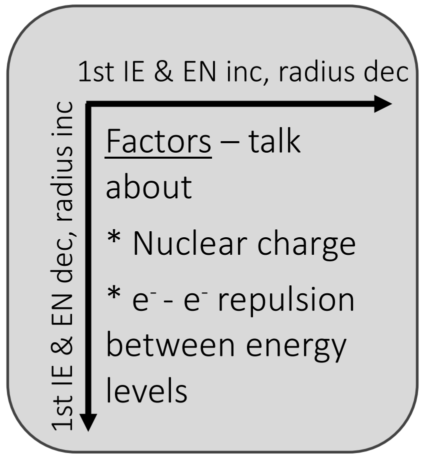
'X' is more electronegative than 'Y', so the X-Y bonds are polar / have dipole.

The dipoles are:

- asymmetrically arranged and therefore do not cancel, so ______ is a polar molecule.
- symmetrically arranged and therefore cancel, so ______ is a non polar molecule.

Exo -ΔH, Endo +ΔH

$\Delta_{fus}H$ $X(s) \rightarrow X(l)$ $\Delta_{vap}H$ $X(l) \rightarrow X(g)$ Fusion and vaporisation both require bonds to be broken between molecules. Fusion requires some (S to L) but vaporisation all (L to G), so more energy is needed. Bond breaking = endo, bond making = exo.



1st IE & EN inc, radius dec

1st IE & EN dec, radius inc

Factors – talk about

* Nuclear charge

* e^- - e^- repulsion between energy levels

$$n = cV$$

$$n = \frac{m}{M}$$

HINT: this m is where the thermometer is placed – mass of water / solution

$$q = mc\Delta T$$

In J

$$\Delta_r H° = \frac{-q}{n}$$

Q is only negative if the reaction is exothermic

$$\Delta_r H° = \Sigma\Delta_f H°(\text{products}) - \Sigma\Delta_f H°(\text{reactants})$$

Data must be f

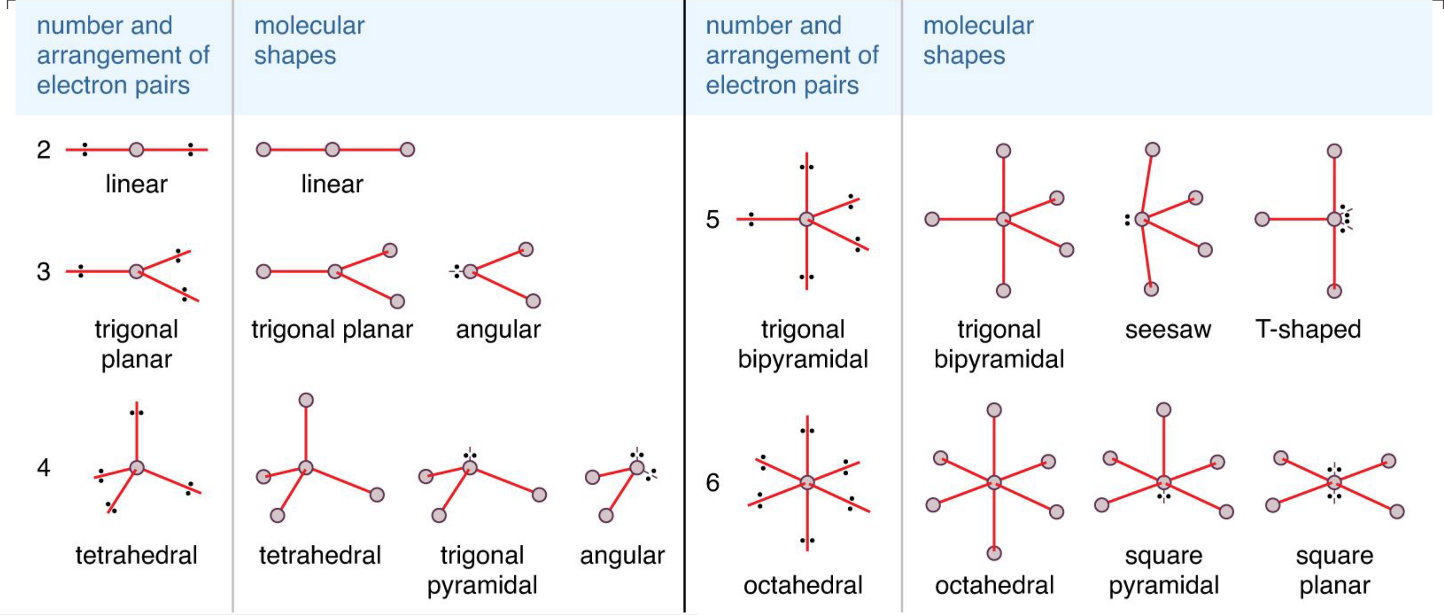
Atomic number / Relative atomic mass

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1 H 1.0																	2 He 4.0
3 Li 6.9	4 Be 9.0											5 B 10.8	6 C 12.0	7 N 14.0	8 O 16.0	9 F 19.0	10 Ne 20.2
11 Na 23.0	12 Mg 24.3											13 Al 27.0	14 Si 28.1	15 P 31.0	16 S 32.1	17 Cl 35.5	18 Ar 40.0
19 K 39.1	20 Ca 40.1	21 Sc 45.0	22 Ti 47.9	23 V 50.9	24 Cr 52.0	25 Mn 54.9	26 Fe 55.9	27 Co 58.9	28 Ni 58.7	29 Cu 63.6	30 Zn 65.4	31 Ga 69.7	32 Ge 72.6	33 As 74.9	34 Se 79.0	35 Br 79.9	36 Kr 83.8
37 Rb 85.5	38 Sr 87.6	39 Y 88.9	40 Zr 91.2	41 Nb 92.9	42 Mo 95.9	43 Tc 98.9	44 Ru 101	45 Rh 103	46 Pd 106	47 Ag 108	48 Cd 112	49 In 115	50 Sn 119	51 Sb 122	52 Te 128	53 I 127	54 Xe 131

Br: $[Ar]\,3d^{10}4s^24p^5$

V: $[Ar]\,3d^34s^2$

Ni^{2+}: $[Ar]\,3d^8$



Standard enthalpy of formation – enthalpy change when 1 mol of substance is formed from its elements in their standard states.

Standard enthalpy of combustion enthalpy change when 1 mol of substance is completely burned in oxygen.

Reactants & products in their standard state – *their state at 25°C*

For IONS – add / subtract electron(s) AND put [ ] brackets around ion and charge outside.

More energy is released forming $H_2O(l)$ rather than $H_2O(g)$ because MORE intermolecular attractions are formed – so reaction is more exothermic.

Lewis structure of ClF_4^-