

Electron Configuration

Electron Orbitals

STOP AND CHECK (PAGE 8)

- Electron orbitals represent the specific space around the atom where an electron is most likely to be whereas electron shell configurations from Level 2 Chemistry do not restrict electrons to a specific place around the atom and show them moving freely around the atom.
- s-orbitals, p-orbitals and d-orbitals.
- The 1st electron shell has s-orbitals, the 2nd electron shell has s-orbitals and p-orbitals and the 3rd electron shell has s-orbitals, p-orbitals and d-orbitals.
- Each individual electron orbital contains two electrons.

Writing Electron Configuration

STOP AND CHECK (PAGE 10)

- $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10}$
- The electron configuration of an atom can be determined by calculating the number of electrons which is equal to the element's atomic number and placing all of the electrons into orbitals, starting with the 1s orbitals and working upwards until all electrons are placed.
- The electron configuration of an anion can be determined by calculating the number of electrons which is equal to the element's atomic number plus one for each charge on the ion and placing all of the electrons into orbitals, starting with the 1s orbitals and working upwards until all electrons are placed. The electron configuration of a cation is determined by first finding the electron configuration of the atom it is made from and then removing one

electron for each charge on the ion. Electrons are removed in the reverse order that they were added. The only exception to this rule is that the 4s orbitals are emptied before any of the 3d orbitals despite the 3d orbitals being filled after the 4s orbitals.

Exceptions to the Rule

STOP AND CHECK (PAGE 11)

- Chromium (Cr) and copper (Cu)
- These are exceptions because it is more stable to have half-filled orbitals and full orbitals as opposed to partially filled orbitals.

Electron Configuration

QUICK QUESTIONS (PAGE 11)

- Calcium (Ca): $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$ OR $[\text{Ar}] 4s^2$
- Iron (Fe): $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$ OR $[\text{Ar}] 4s^2 3d^6$
- Oxygen (O): $1s^2 2s^2 2p^4$
- Iron(III) ion (Fe^{3+}): $1s^2 2s^2 2p^6 3s^2 3p^6 3d^5$ OR $[\text{Ar}] 3d^5$
- Sulphide (S^{2-}): $1s^2 2s^2 2p^6 3s^2 3p^6$ OR $[\text{Ar}]$

Atomic Radius

Trends in Atomic Radius Across a Period

STOP AND CHECK (PAGE 14)

- Across a period atomic radius decreases.
- Across a period nuclear charge increases.
- Despite the increase in repulsion between electrons in the valence shell, the increase in nuclear charge and therefore increase in attraction for the valence electrons by the nucleus is stronger than the repulsion between electrons in the valence shells so radius decreases.

Trend in Atomic Radius within a Periodic Table Group

STOP AND CHECK (PAGE 16)

- Down a group atomic radius increases.
- Down a group the number of electron shells increases.
- The more electron shells there are, the more inner electrons there are between the valence electrons and the nucleus which also experience a force of attraction with the nucleus. These inner electrons shield the valence electrons from the attractive forces of the nucleus. The less the valence electrons are attracted to the nucleus, the further apart they will be.

Atomic Radius

QUICK QUESTIONS (PAGE 16)

- Lithium (Li) and fluorine (F) are both in the 2nd period of the periodic table. Across a period atomic radius decreases as nuclear charge increases. This is because atoms with a higher nuclear charge experience a greater force of attraction on their valence electrons, resulting in the valence electrons being held more closely to the nucleus of the atom, effectively decreasing its radius. Because Li has 3 protons and F has 9 protons, F has a higher nuclear charge and will therefore have a smaller radius than Li.
- Oxygen (O) and sulphur (S) are both in the 16th group of the periodic table. Down a group atomic radius increases as the number of electron shells increases. This is because inner electrons shield the valence electrons from the attractive forces of the nucleus. The more electron shells there are the more shielding is and the smaller the attractive force of the nucleus experienced by the valence electrons, causing the valence electrons to exist further from the nucleus, increasing atomic radius. Because O is in the 2nd period and S is in the 3rd period of the periodic table and further down in group 16, it, therefore, has more electron shells which result in increased shielding and a larger atomic radius.

Ionic Radius

Radius of Cations

STOP AND CHECK (PAGE 17)

- Cations will have a smaller radius than that of the atom they are formed from.
- Metal atoms have a small number of electrons in their valence shell and are able to easily lose them to form cations. This causes the next inner electron shell of the atom to become the valence shell, decreasing the shielding between the nucleus and the valence electrons. This means the valence electrons are held more closely to the nucleus, decreasing the radius of the cation comparatively to the atom it was made from.

Radius of Anions

STOP AND CHECK (PAGE 8)

- Anions will have a larger radius than that of the atom they are formed from.
- The only factor that affects atomic radius that changes when non-metals form anions is the number of electrons in the valence shell. This is because the number of protons remains unchanged and therefore nuclear charge remains the same. Anions are formed when nonmetals gain extra electrons to fill the empty orbitals in their valence shell, but the number of electron shells does not increase, and therefore shielding doesn't increase either.

Ionic Radius

QUICK QUESTIONS (PAGE 19)

- Fluorine (F) gains one electron to fill its 2p electron orbital to form fluoride (F^-). Increasing the number of electrons in the valence shell increases the degree of repulsion between the valence electrons. As the number of protons does not change between the atom and ion, the same nuclear charge is being applied to a greater number of electrons in F^- as opposed to F. This results in the radius

of F^- being larger than that of F due to the increased repulsion of the valence electrons.

- Calcium (Ca) loses two electrons to empty its 4s electron orbital to form the calcium ion (Ca^{2+}). This means that the 3p orbital is the valence orbital of Ca^{2+} . As the 3p orbital is lower energy than the 4s orbital Ca^{2+} has decreased shielding between its valence electrons and its nucleus in comparison with Ca. Because of the decreased shielding, Ca^{2+} has a smaller radius than Ca.

Electronegativity and Ionisation Energy

Ionisation Energy

STOP AND CHECK (PAGE 21)

- Ionisation energy is the energy required to remove one mole of electrons from one mole of atoms in their gaseous state.
- Atoms must be in their gaseous state, as atoms in their solid or liquid state will first use energy to change states from solid to liquid and liquid to gas before any additional energy is used to remove electrons. Therefore, it is important that the atoms are in their gaseous state so that any additional energy used in state changes isn't measured.

Electronegativity

STOP AND CHECK (PAGE 21)

- Electronegativity is the measure of how strongly an atom is attracted to the bonding electrons in a covalent bond.
- Across a period of the periodic table electronegativity increases.
- Down a group of the periodic table electronegativity decreases.
- Electronegativity increases across a period as nuclear charge increases allowing the nucleus to have a stronger force of attraction for bonding electrons. Therefore, bonding electrons will be held to the nucleus more tightly. Electronegativity decreases down a group as shielding increases and the valence electrons are further away meaning the nucleus has a weaker force of

attraction for bonding electrons. Therefore, bonding electrons will be held to the nucleus less tightly.

Electronegativity and Ionisation Energy

QUICK QUESTIONS (PAGE 22)

- Fluorine (F) is further along the 2nd period than carbon (C) on the periodic table. This means that F has a greater nuclear charge which results in a greater force of attraction for bonding electrons. As electronegativity is the measure of how strongly an atom is attracted to bonding electrons in a covalent bond, F will have a greater electronegativity than C as it has a greater force of attraction between its nucleus and bonding electrons.
- Caesium (Cs) is near the bottom of its group resulting in a large number of inner electron shells and therefore a high shielding effect between the attractive forces of the nucleus and its bonding electrons. Cs is also the first element in its period, meaning that it has the smallest nuclear charge of any element that has an equivalent shielding effect. Due to its high shielding and low nuclear charge Cs has one of the smallest values for electronegativity.
- Oxygen (O) is above sulphur (S) in the 16th group of the periodic table. Because of this O has fewer electron shells than S and therefore decreased shielding. Decreased shielding means that O has a stronger force of attraction between its nucleus and its valence electrons than S. Because O holds onto its valence electrons more tightly than S, it takes more energy to remove one of those electrons than it does for S. Therefore, O has higher ionisation energy than S as ionisation energy is the energy required to remove one mole of electrons from one mole atoms in their gaseous state.
- The sodium ion (Na^+) has the same number of protons and therefore nuclear charge as the sodium atom (Na) but has one less electron. Because of this Na^+ has less electron-electron repulsion than Na and therefore a greater electrostatic force of attraction between its nucleus and valence electrons. More energy is required to overcome this greater electrostatic force of attraction to remove an electron from Na^+ compared to Na, meaning that the ionisation energy of Na is smaller than it is for Na^+ .
- Nitrogen (N) has higher ionisation energy than oxygen (O) despite O being further along the 2nd period than N. This is because N has the electron configuration $1s^2 2s^2 2p^3$ and O has the electron configuration $1s^2 2s^2 2p^4$. As the

2p orbital can have up to 6 electrons, the 2p orbital of N is half-filled and the 2p orbital of O is partially filled. Because it is more stable for an atom to have a half-filled or full electron orbital, O is more willing to lose an electron to have a stable half-filled 2p orbital than N is willing to lose an electron and go from a stable half-filled 2p orbital to an unstable partially filled 2p orbital. Because of this, the ionisation energy of N is higher than it is for O, despite O being further along the same period.

Lewis Structures

Expanding the Octet

STOP AND CHECK (PAGE 26)

- Atoms in the 2nd period of the periodic table such as carbon, nitrogen and oxygen are unable to disobey the Octet Rule. This is because they are unable to utilise nearby d-orbitals.
- Atoms in the 3rd period of the periodic table and beyond such as sulphur, phosphorus, chlorine and iodine are able to disobey the Octet Rule. This is because they are able to utilise nearby d-orbitals.

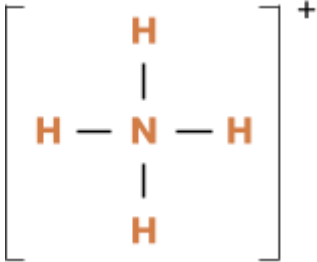
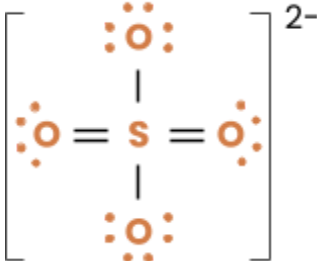
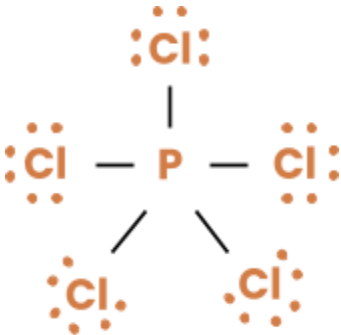
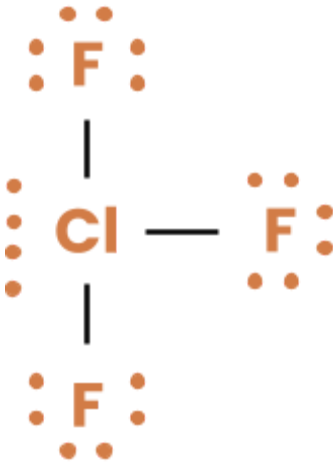
Lewis Structures of Polyatomic Atoms

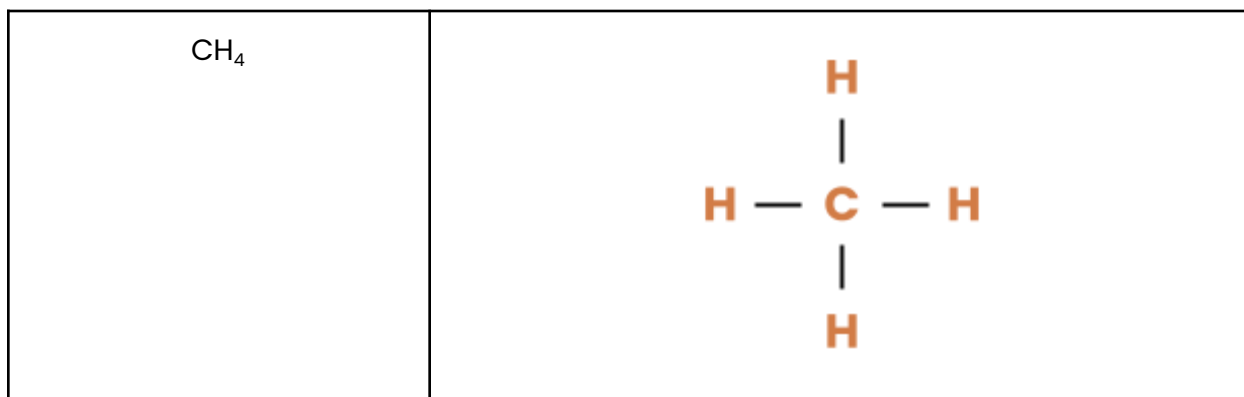
STOP AND CHECK (PAGE 28)

- The Lewis Structures of charged molecules will have extra or fewer electrons depending on the charge of the ion. The Lewis Structures of charged molecules are also drawn in square brackets with the charge of the molecule indicated in the top right-hand corner outside the brackets.

Lewis Structures

QUICK QUESTIONS (PAGE 29)

Chemical Formula	Lewis Structure
NH_4^+	
SO_4^{2-}	
PCl_5	
ClF_3	



Molecular Shapes and Polarity

VSEPR Theory

STOP AND CHECK (PAGE 31)

- Electrons repel each other as they have the same negative charge. Different numbers of electrons repel each other in different shapes which can be predicted by VSEPR.
- The two types of electron pairs around an atom in a molecule are bonding pairs shared between two atoms in a covalent bond and non-bonding or lone pairs.

Polarity

STOP AND CHECK (PAGE 44)

- Electronegativity of atoms within a covalent bond and molecular shape can help to determine polarity.
- Symmetry is important when it comes to polarity as symmetrical molecules tend to always be non-polar as long as all of the atoms around the central atom are the same. This is because the dipoles of the covalent bonds are arranged evenly around the central atom, so their effects cancel out.
- Molecules with more than one type of atom around the central atom are always polar regardless of molecular shape due to differences in

electronegativity between the surrounding atoms, meaning that the dipoles of the covalent bond are not equivalent and will not cancel each other out even if the molecule is symmetrical.

Molecular Shapes and Polarity

QUICK QUESTIONS (PAGE 44)

- SF_6 is an octahedral molecule with six polar S-F bonds arranged symmetrically around the central S atom. Because of this, the polar bonds cancel each other out and the overall molecule is nonpolar.
- CH_3Cl is a tetrahedral molecule with three nonpolar C-H bonds and one polar C-Cl bond arranged symmetrically around the central C atom. Because the molecules around the central C atom are not all the same the bond polarities are therefore not equivalent and are unable to cancel each other out despite CH_3Cl being symmetrical and the overall molecule is polar.
- XeF_4 is a square planar molecule with four polar Xe-F bonds arranged symmetrically around the central Xe atom. Because of this, the polar bonds cancel each other out and the overall molecule is nonpolar.
- IF_5 is a square pyramidal molecule with five polar I-F bonds arranged asymmetrically around the central I atom. Because of this, the polar bonds are unable to cancel each other out and the overall molecule is polar.

Intermolecular Forces

Instantaneous Dipole Forces

STOP AND CHECK (PAGE 46)

- All molecules are held together by instantaneous dipole forces.
- As molecule size increases, the number of electrons within the molecule also increases. The more electrons there are the stronger the instantaneous positive and negative charges are and the more often they occur resulting in stronger instantaneous forces of attraction.

- Nonpolar or uncharged molecules are able to attract one another by inducing instantaneous dipoles in other nearby molecules as high concentrations of electrons in a negative dipole will repel electrons in other molecules creating a positive dipole. As opposite charges attract, the positive instantaneous dipoles of molecules are attracted to the negative instantaneous dipoles of other molecules.

Permanent Dipole Forces

STOP AND CHECK (PAGE 48)

- Polar molecules are held together by permanent dipole forces.
- Permanent dipole forces exist between polar molecules due to opposite charges attracting one another between the positive and negative ends of the molecules. Temporary dipole forces exist between all molecules, regardless of polarity due to the random movement of electrons, creating temporary dipoles in areas with high or low electron density.

Hydrogen Bonding

STOP AND CHECK (PAGE 49)

- A molecule needs to have a hydrogen atom covalently bonded to either a nitrogen, oxygen or fluorine atom for it to be able to hydrogen bond.
- Hydrogen bonding also involves polar bonds like with permanent dipole forces, but there is a much greater difference in electronegativities between hydrogen and either nitrogen, oxygen or fluorine. Because the bond dipoles in hydrogen bonding are more positive or negative than the bond dipoles in permanent dipole forces, hydrogen bonds are stronger.

Melting and Boiling Points

STOP AND CHECK (PAGE 50)

- At their melting point, solids have enough heat energy to break some of the bonds between molecules so it becomes a liquid. At their boiling point, liquids

have enough heat energy to break all of the bonds between molecules so it becomes a gas.

- The melting and boiling points of molecules can be used to determine the strength of the intermolecular forces.
- Temporary dipole forces are the weakest intermolecular forces and hydrogen bonds are the strongest intermolecular forces.

Intermolecular Forces

QUICK QUESTIONS (PAGE 51)

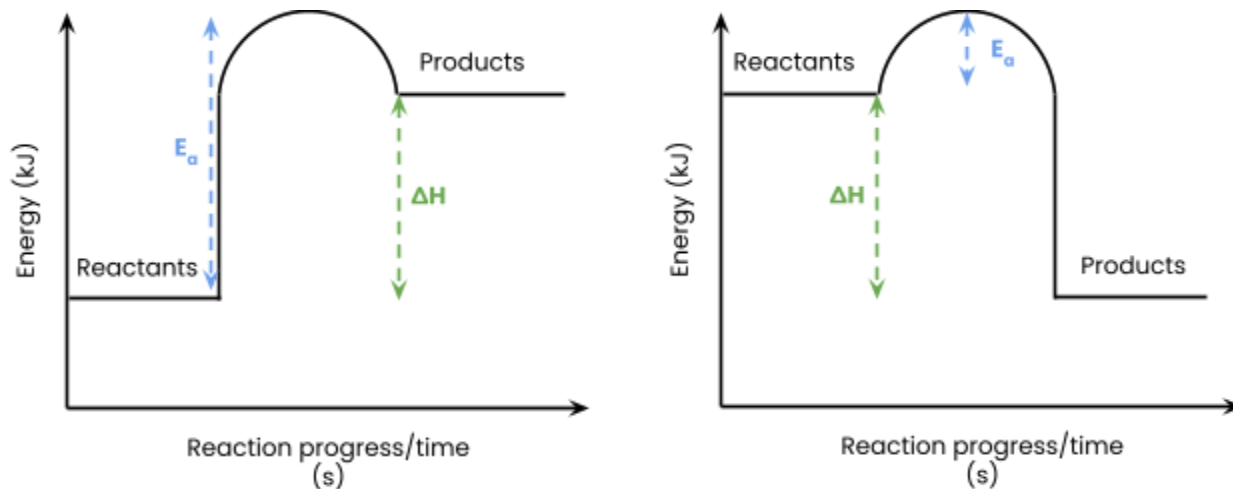
- Nitrogen has the lowest melting/boiling point as it is a nonpolar molecule and only has temporary dipole forces between the molecules. Nitric oxide has a higher melting/boiling point than nitrogen as it is a polar molecule and has permanent dipole forces between the molecules which are stronger than temporary dipole forces. Ammonia has a higher melting/boiling than both nitric oxide and nitrogen as it contains N-H hydrogen bonds which are much stronger than permanent and temporary dipole forces.

Thermochemistry

Enthalpy

STOP AND CHECK (PAGE 58)

- In a chemical reaction the system represents the reactants and the products, whereas the surroundings represent everything in the universe that is not a reactant or product in the specified reaction.



- Both endothermic and exothermic reactions have an activation energy which is an amount of energy that must be provided to the system for the reaction to occur.
- Enthalpy changes are measured in reactions done at standard conditions, or at 25°C and 1 atmosphere. At standard conditions the state at which a substance exists is defined as its standard state.
- Definitions:
 - **Enthalpy of formation** is the energy required to produce one mole of product from its constituent atoms in their standard states at standard conditions.
 - **Enthalpy of combustion** is the energy required to react one mole of reactant with oxygen under standard conditions.
 - **Enthalpy of vaporisation** is the energy required to convert one mole of a liquid into one mole of gas at its boiling point.
 - **Enthalpy of fusion** is the energy required to convert one mole of a solid into one mole of liquid at its melting point.
 - **Enthalpy of sublimation** is the energy required to convert one mole of a solid into one mole of gas at its sublimation point.
- Reactions:
 - $2\text{C}_{(\text{s})} + 2\text{H}_{2(\text{g})} + \text{O}_{2(\text{g})} \rightarrow \text{CH}_3\text{COOH}_{(\text{l})}$
 - $\frac{1}{2}\text{N}_{2(\text{g})} + \frac{1}{2}\text{H}_{2(\text{g})} \rightarrow \text{NH}_{3(\text{g})}$
 - $\text{CH}_3\text{CH}_2\text{OH}_{(\text{l})} + 3\text{O}_{2(\text{g})} \rightarrow 2\text{CO}_{2(\text{g})} + 3\text{H}_2\text{O}_{(\text{l})}$
 - $\text{C}_{(\text{s})} + \text{O}_{2(\text{g})} \rightarrow \text{CO}_{2(\text{g})}$
 - $\text{CO}_{2(\text{s})} \rightarrow \text{CO}_{2(\text{g})}$

Enthalpy Calculations Revision

STOP AND CHECK (PAGE 61)

- Number of moles (n) = $\frac{\text{mass (m)}}{\text{molar mass (M)}}$
- Mass (m) = number of moles (n) \times molar mass (M)
- Enthalpy (ΔJ) = number of moles (n) \times enthalpy change of reaction ($\Delta_r H$)
- Enthalpy change of reaction ($\Delta_r H$) = $\frac{\text{enthalpy (}\Delta J\text{)}}{\text{no. of moles (n)}}$

Heat and Specific Heat

STOP AND CHECK (PAGE 63)

- Heat transferred (q) = mass of substance heated (m) \times specific heat capacity of substance (c) \times change in temperature (ΔT)
 - $q = mc\Delta T$
- Specific heat capacity is the amount of heat required to increase the temperature of a compound by 1°C.
- Compounds with a low specific heat capacity require less energy in order to change their temperature and are therefore easier to heat than compounds with a high specific heat capacity.

Entropy

STOP AND CHECK (PAGE 66)

- Entropy describes the randomness or the disorder of a system.
- Entropy increases as you go from solid to liquid to gas. This is because as you change states from solid to liquid or liquid to gas the kinetic energy of the system increases as the molecules move more freely and more quickly. The more random and rapid movement there is within a system, the more disorder and therefore the more entropy there is.
- Entropy increases when more moles of a substance are produced because this increases the disorder of the system. The more moles the more particles there are to move around randomly.

Spontaneous and Nonspontaneous Reactions

STOP AND CHECK (PAGE 68)

- Spontaneous reactions occur independently without any external factors such as a source of energy. Nonspontaneous reactions do not occur independently and are dependent on an external factor such as a source of energy.
- If a reaction is exothermic and increases entropy it will always be spontaneous.
- If a reaction is endothermic and decreases entropy it will always be nonspontaneous.
- It can be difficult to make a call on whether a reaction is spontaneous or not as some reactions have features of both spontaneous and nonspontaneous reactions such as being exothermic and decreasing entropy or being endothermic and increasing entropy.

Hess's Law

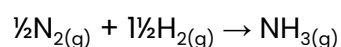
STOP AND CHECK (PAGE 71)

- Hess's Law states the total enthalpy change during a reaction is the same whether the reaction is made in one step or in several steps, made up of multiple component reactions.
- Hess's Law can be used to calculate the enthalpy change of a reaction by rearranging and combining equations with known enthalpy changes so that the final equation is identical to the equation of the reaction that you are trying to calculate the enthalpy for.

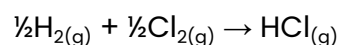
Thermochemistry

QUICK QUESTIONS (PAGE 72)

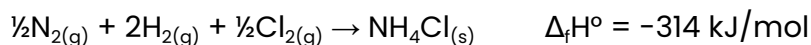
- **Step 1:** Write out the formation reactions provided:



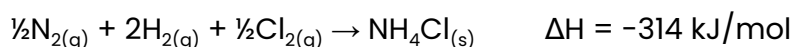
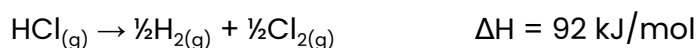
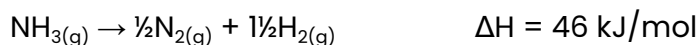
$$\Delta_f H^\circ = -46 \text{ kJ/mol Reverse Equation}$$



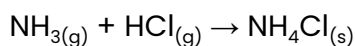
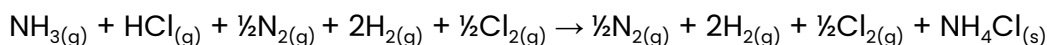
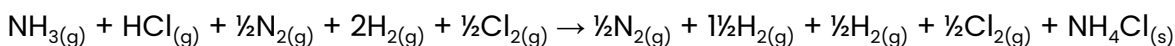
$$\Delta_f H^\circ = -92 \text{ kJ/mol Reverse Equation}$$



Steps 2 and 3: Multiply and/or reverse equations so products and reactants of the unknown enthalpy reaction are on the right side of the equation:

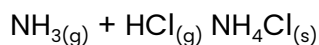


Steps 4 and 5: Combine the given equations and cancel out any molecules that appear on both sides of the reaction:



Net equation is equal to the equation of the reaction of unknown enthalpy.

Step 6: Add together the known enthalpies to calculate the unknown enthalpy:



$$\Delta_r H^\circ = 46 \text{ kJ/mol} + 92 \text{ kJ/mol} + (-314 \text{ kJ/mol})$$

$$\Delta_r H^\circ = -176 \text{ kJ/mol}$$

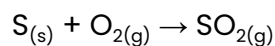
- $q = mc\Delta T$
 - $\Delta T = 33.3^\circ\text{C} - 15.6^\circ\text{C}$
 - $\Delta T = 17.7^\circ\text{C}$
 - $q = 750\text{g} \times 4.18\text{J/g}^\circ\text{C} \times 17.7^\circ\text{C}$
 - $q = 55.4895 \text{ kJ}$

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

- $n = \frac{5.12\text{g}}{32.0\text{g/mol}}$
- $n = 0.16 \text{ mol}$

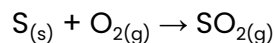
Combustion of methanol ($\Delta_c H^\circ = \frac{\text{heat transferred}}{n}$)

- $\Delta_c H^\circ = \frac{55.4895}{0.16}$
 - $\Delta_c H^\circ = 346.8093 \text{ kJ/mol}$
 - $\Delta_c H^\circ = 347 \text{ kJ/mol}$ (0 d.p.)
- Formation of sulphur dioxide gas:



One mole of sulphur dioxide gas is being formed from its constituent elements at standard states.

Combustion of solid sulphur:



One mole of solid sulphur is being reacted with oxygen gas.

As the equations for both reactions are exactly the same, the enthalpy of formation of sulphur dioxide gas and the enthalpy of combustion of solid sulphur will be exactly the same.