

Module 0: Review

Solutions: a solution is a homogeneous (uniform composition of molecules) mixture of one or more solutes in a solvent. The solvent determines the phase of the solution, it is usually the most abundant component. The solutes are all other components of a solution.

Ionic Compound: Comprised positive cations and negative ions (ions are atoms with a different charge from normal, due to loss or gain of electrons) in repeating patterns. Ions are held together by strong ionic bonding forces. Solid at room temperature, but can be dissolved in water, its ions are separated in the solution (dissociation).

Molecular Compound: Comprised of stable, neutral molecules, with state varying on intermolecular forces. A molecule may or may not produce ions when dissolved in water (ionization).

Electrolyte: ionic and molecular compounds which produce ions when dissolved in water. If they completely dissociate or ionize they are said to be strong, otherwise they are said to be weak.

Concentration: Concentration (C) measured in mol L^{-1} is the number of moles per litre of solution, the solubility (S) is the maximum molar concentration of a solution. Solution is unsaturated if $C < S$, saturated if $C = S$, superaturated if $C > S$.

Solubility Table:

	NH_4^+	Li^+	Na^+	K^+	Mg^{2+}	Ca^{2+}	Ba^{2+}	Al^{3+}	Fe^{3+}	Cu^{2+}	Ag^+	Zn^{2+}	Pb^{2+}
F^-	sol	sol	sol	sol	insol	insol	sl sol	sol	sl sol	sol	sol	sol	insol
Cl^-	sol	sol	sol	sol	sol	sol	sol	sol	sol	sol	insol	sol	sol
Br^-	sol	sol	sol	sol	sol	sol	sol	sol	sol	sol	insol	sol	sl sol
I^-	sol	sol	sol	sol	sol	sol	sol	sol			insol	sol	insol
OH^-	sol	sol	sol	sol	insol	sl sol	sol	insol	insol	insol		insol	insol
S^{2-}	sol	sol	sol	sol		sl sol	sol		insol	insol	insol	insol	insol
SO_4^{2-}	sol	sol	sol	sol	sol	sl sol	insol	sol	sol	sol	sl sol	sol	insol
CO_3^{2-}	sol	sol	sol	sol	insol	insol	insol			sl sol	insol	insol	insol
NO_3^-	sol	sol	sol	sol	sol	sol	sol	sol	sol	sol	sol	sol	sol
PO_4^{3-}	sol	insol	sol	sol	insol	insol	insol	insol	insol	insol	insol	insol	insol
CrO_4^{2-}	sol	sol	sol	sol	sol	sol	insol		insol	insol	insol	insol	insol
CH_3CO_2^-	sol	sol	sol	sol	sol	sol	sol	sl sol	sol	sol	sol	sol	sol

Bronsted-Lowry Acid and Base: An acid is a proton (H^+) donor and a base is a proton (H^+) acceptor. By this definition an acid-base reaction is a proton transfer.

Common Strong Acids and Bases:

Formula	Acid Name	Formula	Base Name
HCl	Hydrochloric acid	$\text{LiOH}, \text{NaOH}, \dots$	Group 1 hydroxides
HBr	Hydrobromic acid	$\text{Mg(OH)}_2, \text{Ca(OH)}_2, \dots$	Group 2 hydroxides
HI	Hydroiodic acid	H^-	Hydride
HClO_4	Perchloric acid	O^{2-}	Oxide
HBrO_4	Perbromic acid		
H_2SO_4	Sulfuric acid		
HNO_3	Nitric acid		

Acid-base Neutralization: Acid and base reacting to form a salt and often water. If the acid or base is strong then the neutralization is complete.

Rules for Assigning Oxidation States: The oxidation state of an atom in a compound is the charge it would have if all electrons went to the most electronegative atom in a bond. General rules in order of priority:

- Oxidation state is zero for atom in elemental form.
- The sum of oxidation states must equal the charge of the compound.
- Group 1 and 2 metals have oxidation state +1 and +2 respectively.
- F always has an oxidation state of -1 .
- H normally has oxidation state +1, except when combined with group 1 or 2 metals.
- O normally has oxidation state -2 except when bonded to itself or when bonded to F .
- $Cl, Br,$ and I have oxidation state of -1 , unless dictated otherwise.

Oxidation-Reduction (redox) Reaction, oxidation states change. The Oxidized atom Is the one which Loses electrons, the Reduced atom Is the one which Gains electrons (think OILRIG). A half reaction is the initial oxidized or reduced atom yielding its final state, balanced with e^- so there are equal electrons on either side. To balance a redox reaction, complete the following steps.

- Determine oxidation state of each atom.
- Write half reactions for each.
 - Balance atoms that change oxidation state.
 - Determine number of electrons gained or lost
 - Balance charges by using H^+ (in acidic solutions) or OH^- (in basic solutions).
 - Balance the rest of the atoms using H_2O .
- Balance the number electrons transferred in each half reaction so they cancel.
- Add the two half reactions together.

Module 1: The Periodic Table

Isotopes and Allotropes: Isotopes: atoms with the same number of protons but different number of neutrons. Allotropes: different molecular forms of the same element, e.g. graphite and diamond are allotropes of carbon.

Atomic Mass Intervals: Atomic mass intervals provide an interval for the weight based on the abundance of certain isotopes. Conventional atomic mass is the most common weight.

Atom Economy: Theoretical percentage of reactant which becomes product. $\%AE = \frac{\text{stoichiometric mass of desired product}}{\text{stoichiometric mass of reactants}} \times 100\%$

E-Factor: Number of kilograms of waste generated per kilogram of product. $\text{E-Factor} = \frac{\text{mass of non-water waste produced}}{\text{mass of product obtained}} \times 100\%$

Module 2: Quantum Theory and the H Atom

Wavelength-Frequency Relation: Wavelength, λ (m): distance between maxima. Period, T (s): time between maxima. Frequency, $\nu = \frac{1}{T}$ (s^{-1}): number of times the wave reaches a maximum in a second. $\lambda\nu = c$, where $c = 2.998 \cdot 10^8 \text{ m s}^{-1}$ is the speed of light.

Electromagnetic Spectrum:

Name	Wavelength (λ), m	Frequency (ν), s^{-1}	Energy (E), eV
Gamma-ray	10^{15}	$3 \cdot 10^{23}$	$1.2 \cdot 10^9$
X-ray	10^{-9}	$3 \cdot 10^{17}$	$1.2 \cdot 10^5$
Ultraviolet	10^{-7}	$3 \cdot 10^{15}$	$1.2 \cdot 10^3$
Visible (Blue)	$4 \cdot 10^{-7}$	$7.5 \cdot 10^{14}$	3.1
Visible (Red)	$7 \cdot 10^{-7}$	$4.3 \cdot 10^{14}$	1.8
Infrared	10^{-5}	$3 \cdot 10^{13}$	0.12
Microwave	0.01	$3 \cdot 10^{10}$	$1.2 \cdot 10^{-4}$
Radio wave	100	$3 \cdot 10^7$	$1.2 \cdot 10^{-8}$

Photoelectric Effect: Light used to dislodge electrons from the surface of a metal. Maximum kinetic energy (KE) is measured as a function of frequency (ν) and intensity (I) of the light. Electrons are ejected if and only if ν is greater than some threshold ν_0 . If $\nu > \nu_0$, the KE of an ejected electron increases proportionally to ν , and electrons are ejected instantaneously.

Einstein Explanation of the Photoelectric Effect: Light must be concentrated in particle-like bundles, photons. The energy of a photo is proportional to the frequency of the light, $E_{\text{photon}} = h\nu$ where $h = 6.626 \cdot 10^{-34} \text{ J s}$ is Planck's constant. Alternatively, $E_{\text{photon}} = w + (KE)_{e^-} = h(\nu - \nu_0)$ where w is the work function of the metal, the threshold amount of energy to dislodge an electron.

Bohr Model of the H Atom: Electrons orbit energy levels around the atom and when they go from one energy level to another they absorb (positive) or release (negative) energy equal to $E_{\text{upper}} - E_{\text{lower}}$ where E_{upper} and E_{lower} are the energies at the respective energy levels. He derived for formula for n th energy level (with $n = 1$ as innermost) to be $E_n = -\frac{R_H}{n^2}$ where $R_H = 2.179 \cdot 10^{-18} \text{ J}$ is the Rydberg constant.

de Broglie's Hypothesis: All matter posses wave-particle duality. Any particle with mass m travelling with speed v possesses wave length $\lambda_{dB} = \frac{h}{mv}$.

Heisenberg Uncertainty Principle: $\Delta x \Delta p \geq \frac{h}{4\pi}$ where Δx is the uncertainty in the position of the particle and $\Delta p = m\Delta v$ is the uncertainty in the momentum of the particle. It is impossible to know both the position and momentum of a particle with absolute certainty.

Quantum Numbers: Note ℓ is also called the angular momentum.

Name	Symbol	Range	Property
Principal	n	1, 2, 3, ...	Orbital size and energy
Secondary	ℓ	0, 1, ..., (n-1)	Orbital shape (sub-shell)
Magnetic	m_ℓ	$-\ell, -\ell+1, \dots, \ell-1, \ell$	Orbital orientation
Spin	s	only $\frac{1}{2}$	Electron spin magnitude
Magnetic spin	m_s	$\frac{1}{2}$ and $-\frac{1}{2}$	Electron spin direction

ℓ is associated with particular sub-shells: $\frac{\ell}{\text{Sub-shell}} \quad \begin{matrix} \ell & 0 & 1 & 2 & 3 \\ \text{Sub-shell} & s & p & d & f \end{matrix}$

Schrödinger's Results for H Atom: Each stable state of the H atom gives a function Ψ called the orbital. $|\Psi|^2$ gives the probability of finding an electron at any particular point. Discovery of quantum numbers (see above). Each stable state of the H atom has well-defined energy $E_n = -\frac{R_H}{n^2}$.

H Atom Wave Functions: Each orbital satisfies an equation involving a radial factor $R(r)$ and angular factor $Y(\theta, \phi)$ in polar co-ordinates, $\Psi = R(r)Y(\theta, \phi)$. Most orbitals have radial nodes (circular) where $R = 0$ and angular nodes (linear) where Y changes signs, these are places with 0 probability of finding electrons. An orbital has $n - \ell - 1$ radial nodes and ℓ angular nodes.

Representations of H Atom Wave Functions:

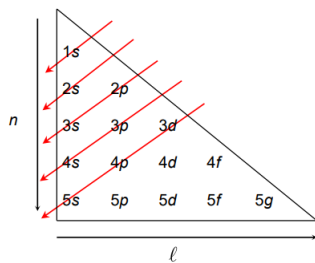
- Scatter plots: empty circles for radial nodes and empty lines for angular nodes. Tells us about distribution of electrons.
- Boundary surface plots: 3d models of regions of space with high (usually 95%) probability of finding electrons. Tells us about the shape and orientation of electron clouds. Nodes where there is empty space.
- Radial factors: plots of $R(r)$. $R(r)$ decreases exponentially as r increases. $R(r)$ crosses r axis as many times as there are radial nodes. s sub-shells have maximum value at $r = 0$, other sub-shells have $R(0) = 0$.
- Radial electron densities: plots $R(r)^2$ vs r . Maxima progressively decrease as r increases. Has one more maxima than number of radial nodes. s sub-shells have maximum value at $r = 0$, other sub-shells have $R(0)^2 = 0$. Tells us the density of an electron cloud or probability of finding an electron as we go in a certain direction (so do radial factors).
- Radial distribution plots: plots of $r^2R(r)^2$ vs r . Attains maximum value at most probable distance between electron and nucleus, r_{mp} . Maxima progressively increase as r increases. All plots start at 0. r_{mp} increases as n increases. For fixed n , r_{mp} decreases as ℓ increases.

Module 3: Multielectron Atoms

Penetration and Shielding: Penetration is the measure of how close a particular electron gets to the nucleus. Shielding is a measure of how thoroughly blocked (shielded) an electron is from the nucleus. As the atomic number Z increases, the attraction between the nucleus and any electron increases, lowering the energy of the orbital.

Ground State Electron Configuration: To minimize the energy of the atom, electron distribute themselves according to three rules:

1. Pauli Exclusion Principle: no two electrons can have the same set of quantum numbers n, ℓ, m_ℓ, m_s . Implies each ℓ sub-shell has $2\ell + 1$ orbitals and $4\ell + 2$ total electrons.
2. Aufbau Procedure: Neutral orbitals are filled minimizing $n + \ell$, with the exception of Cr and Cu which fill s sub-shells before p sub-shells.
3. Hund's Rule: The lowest energy arrangement is the one with the most parallel spins.



Model of the Aufbau procedure in practice.

Periodic Table: Each of the 7 row of the periodic table, called periods, starts a shell. Each of the 18 columns of the periodic table, called groups, have similar electron configurations. An atom has a number of valence (outermost shell) electrons equal to $\begin{cases} \text{group \#} & \text{if group \#} < 12 \\ \text{group \#} - 10 & \text{if group \#} \geq 12 \end{cases}$

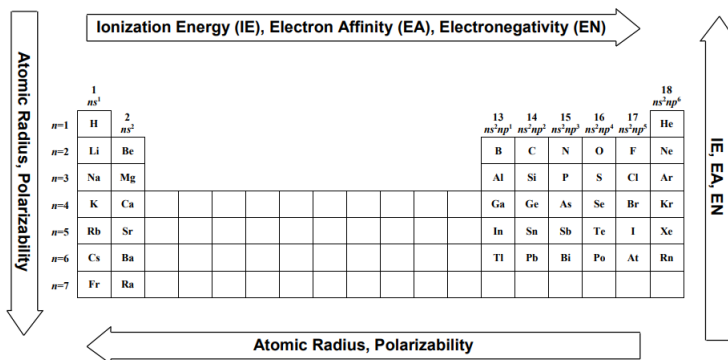
Determining Electron Configuration of Monatomic Ions: Determine the ground state of the neutral atom first. Then if a negative charge (anion) add e^- s according to the $n + \ell$ rule. If positive charge (cation) remove e^- s with highest n value first, if equal then highest ℓ first.

Magnetism: An atom is diamagnetic if and only if all electrons are paired. In this case the atom does not possess a magnetic moment and only interacts weakly with external magnetic fields. An atom is paramagnetic if and only if there is one or more unpaired electrons. In this case the atom possesses a magnetic moment and interacts strongly with external magnetic fields.

Atomic Radius: The distance between nuclei in a certain environment. The covalent radius is $\frac{1}{2}$ the diatomic bond length the in X_2 molecule. Metallic radius $\frac{1}{2}$ of the radius between nearest neighbours in a metallic solid. Atomic radius decreases down a period and increases down a group.

Ionization Energy: Energy required to remove an electron from gas-phase atom (often reported as an enthalpy change ΔH). It becomes increasingly difficult to remove more electrons. Ionization energy decreases down a group and generally increases across as period except when going from group 2 to 13 and group 15 to 16. This decrease is due to electrons sharing shells, either because of repulsion or shielding.

Electron Affinity: Energy change from the addition of an electron to a gas-phase atom. Most atoms (except N and those in groups 2 and 18) release energy when they acquire electrons (negative). Electron affinity tends to increase going across a period, except for groups 2 and 18, and decrease going down a group.



Module 4: Introduction to Chemical Bonding

Compound Guidelines: Metals from s block combine with atoms from p block to form binary, (usually) ionic compounds with ionic bonds. Atoms from p block with each other form (usually) molecular compounds with covalent bonds.

Octet Rule: Atoms exhibit the tendency to attain noble gas configuration by sharing or transferring electrons.

Formal Charge: Number of valence electrons of an atom minus the number of valence electrons it would receive if electrons in each bond were shared equally (formal charge = valence e^- - dots - lines in Lewis structures). Sum of formal charges must equal the total charge of the compound.

Oxidation States vs Formal Charge: Oxidation state is the hypothetical charge if all bonding electrons go to most electronegative atom. Oxidation state exaggerates ionic character of bonding. Oxidation states are used to predict and rationalize chemical properties

Formal charge is the hypothetical charge if all bonding electrons are shared equally. Formal charges exaggerate covalent character of bonding. Formal charges are used to assess the best LEwis structure.

Lewis Symbols and Lewis Structures: The Lewis symbol of an atom is the atomic symbol surrounded by the correct number of valence electrons. Lewis structures (rarely done for ionic compounds) are groups of Lewis symbols with bonds. To draw the Lewis structure of a molecular compounds:

1. Use group numbers to determine number of valence electrons.
2. Put a pair of electrons between each pair of bonded atoms.
3. Distribute remaining atoms keeping in mind
 - An H atom never has more than 2 valence electrons.
 - 2nd period atoms never have more than 8 valence electrons.
 - Atoms from 3rd period and beyond may have expanded octets with more than 8 valence electrons.
4. Assign a formal charge to each atom

Resonance Structure: Compounds with same spatial arrangement of atoms but different distributions of electrons. Structures with negative formal charges on electronegative elements are favoured. Structures that minimize the number of formal charges with formal charges between -1 and 1 are favoured. Structures with the same formal charges on adjacent atoms are highly unfavoured.

Isomeric Structure: Isomers are compounds with the same molecular formula but different spatial arrangements. Isomers are different compounds with different properties.

Lattice Energy: Energy released when gas phase ions combine to form an ionic solid. The greater the charges on the ions, the greater the lattice energy, and the smaller the ionic radii, the greater the lattice energy.

Strength of Covalent Bonds (Bond Energy): The strength of a covalent bond depends on the size of the atoms the bond-order and the polarity of the bond. Bond dissociation energy is the energy required to break a particular covalent bond. Small atoms tend to form strong compact bonds, and as bond length decreases and bond strength increases, bond order increases. Bond energies are approximates because they are averaged, but they're useful because they can estimate $\Delta_r H$ for gas-phase reactions. In particular
$$\Delta_r H = \sum_{\text{reactants}} BDE'_s - \sum_{\text{products}} BDE'_s.$$
















Electronegativity (EN): Quantitative measure of the pull an atom has on electrons in its bonds.

Polar Covalent Bond (Dipole Moment): Uneven distribution of electrons. I.e. a bond has a slightly positive ($+\delta$) and negative ($-\delta$) ends. Dipoles don't indicate true formal charge. A molecule possesses a dipole moment (μ) if the centre of positive charge is not equal to the centre of negative charge. A diatomic molecule has dipole moment $\mu = |\delta|r$ where δ

is the charge of the dipoles and r is the distance from its positive dipole to negative dipole. Dipole decreases as the size of the molecule increases.

Electrostatic Map: Red represents high electron density $-\delta$, blue represents low electron density $+\delta$.

VSEPR Theory: Imperfect but useful model of what atoms and their bonds look like in 3-D space.

Number of electron pairs	Electron pair geometries: 0 lone pair	1 lone pair	2 lone pairs	3 lone pairs	4 lone pairs
2	 Linear				
3	 Trigonal planar	 Bent or angular			
4	 Tetrahedral	 Trigonal pyramidal	 Bent or angular		
5	 Trigonal bipyramidal	 Sawhorse or seesaw	 T-shape	 Linear	
6	 Octahedral	 Square pyramidal	 Square planar	 T-shape	 Linear

Determine if Molecule is Polar: Non-polar if and only if VSEPR model of molecule is symmetrical along each bond.

Module 5: VB and MO Theories

Valence Bond (VB) Theory: Bonding is represented by overlapping atomic orbitals.

σ Bond: Orbitals overlap on Inter-Nuclear Axis (INA) single bond is always a σ bond. More effective than π bond.

π Bond: Orbitals overlap "off-axis," (not on INA) double/triple bonds always have a π bond. In a double bond, one of the bonds is σ and the other is π , in a triple bond there are two π bonds.

Hybridization: Combining atomic orbitals of central to model overall molecule orbital. Approximate by adding and subtracting electron density from orbitals.

VSEPR to Hybridization: The number of groups in the VSEPR model of a molecule should match the number of hybridized orbitals. d orbitals do not participate in bonding, leave them as lone pairs. n atomic orbitals \Rightarrow n hybrid orbitals.

Hybrid orbitals always form σ bonds.

# of groups	Hybridization	Shape	Angle
2	sp	linear	180°
3	sp^2	trigonal planar	120°
4	sp^3	tetrahedral	109.5°
5	sp^3d	trigonal bipyramidal	varies
6	sp^3d^2	octahedral	90°

Hybridization in Multiple Bonds: If a central atom is double or triple bonded, hybridize all of its σ bonds, the remaining bonds should be manifested as p to p bonds, perpendicular to INA.

Determine Hybridization Scheme:

1. Write the Lewis structure.
2. Determine VSEPR geometry of central atom.
3. Identify the conforming hybridization scheme.
4. Sketch bonding orbitals of central and terminal atoms.

Molecular Orbital (MO) Theory: Molecular orbitals are obtained by solving Schrödinger's equation, but approximated as linear combinations of atomic orbitals. Two atomic orbitals (AO) yield two molecular orbitals (MO).

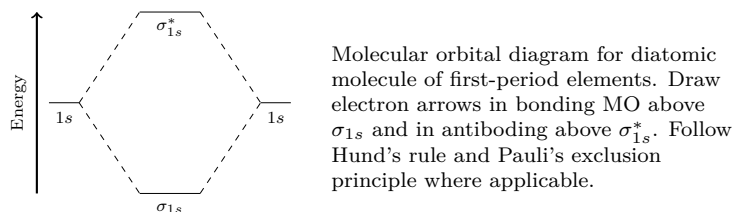
Bonding MO's: A bonding MO is lower in energy than the AO from which it is derived, it is antibonding if it is higher in energy.

Bond Order: defined as

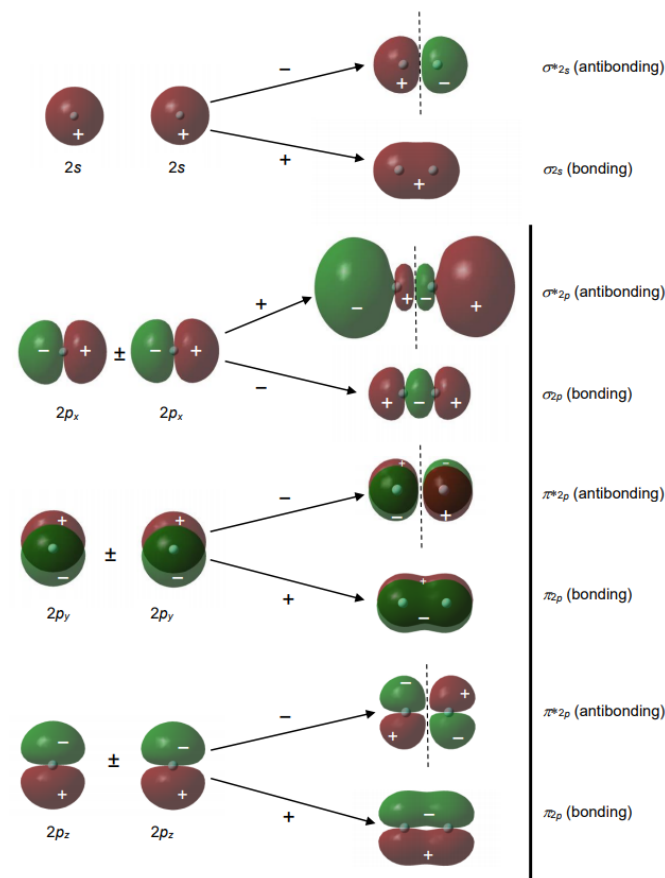
$$\text{bond order} = \frac{\#e^- \text{ bonding orbitals} - \#e^- \text{ antibonding orbitals}}{2}$$

If bond order is positive than the molecule is stable and exists in nature, otherwise it is unstable and doesn't. Higher bond order yields stronger and therefore shorter bonds.

1s MOs: Can obtain an MO either by subtracting or adding the two atomic orbitals. ($\Psi_{tot} = \phi_{1sA} \pm \phi_{1sB}$). Adding produces a bonding MO, with high electron density at A and at B , with moderate in between (σ_{1s}). Subtracting produce antibonding MO, with high electron density at A and at B , with none in between, called a nodal plane (σ_{1s}^*).

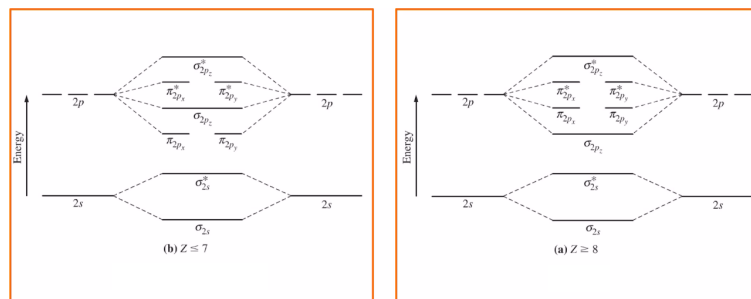


MOs in Second Period:



Note the $1s$ and $2s$ MOs look the same in shape and only differ in size. Note that antibonding MOs always have a nodal plane along the INA and bonding MOs never have one. We denote σ_{2p_x} , π_{2p_y} , π_{2p_z} to specify the added AOs. We implicitly define the INA to be along the axis of the σ_{2p} bond.

Filling MO Energy Diagrams in Second Period: Follows Hund's rule and Pauli's exclusion principle, filling lowest energy MOs first then preferring unpaired electrons. If the atomic number of the atom is greater than 7, the σ_{2p_z} orbital is lower in energy than the π_{2p_x} , π_{2p_y} orbitals.



Module 6: Ideal Gases, Real Gases, and Intermolecular Forces

Kinetic Molecular Theory: Six main assumptions:

1. The molecules in a gas are small and very far apart (most of the volume a gas occupies is empty space).
2. The molecules are in constant random motion (just as many molecules are moving in one direction as in any other).
3. Molecules collide with each other and with the walls of the container (collisions with the walls cause gas pressure).
4. When collisions occur, kinetic energy is conserved i.e. the collisions are perfectly elastic (total energy is constant).
5. The molecules exert no attractive or repulsive forces on each other except in collision, i.e. between collisions they move in straight lines.
6. The average translational kinetic energy of the molecules is proportional to the temperature (k). In particular, $KE_{avg} = (3/2)RT$.

Ideal Gas Law: Denote P (kPa) : pressure, V (L) : volume, T (K) : temperature, $R = 8.3145 \frac{\text{kJ}}{\text{K mol}}$, n (mol) : amount of gas, then $PV = nRT$.

Charles' Law: $V \propto T \iff \frac{V}{T} = \text{const} \iff \uparrow V \uparrow T$.

Boyle's Law: $P \propto \frac{1}{V} \iff PV = \text{const} \iff \uparrow P \downarrow V$.

Avogadro's Law: $V \propto n \iff \frac{V}{n} = \text{const} \iff \uparrow V \uparrow n$.

Gas Mixtures: Suppose a mixture of gases "A + B + C + ..." exists with n_A, n_B, n_C, \dots moles of each. Define $x_A = \frac{n_A}{n_{tot}}$ to be the molar fraction and $P_A = x_A P_{total}$ to be the partial pressure. Note $n_A + n_B + \dots = n_{tot}$, $x_A + x_B + \dots = 1$, $P_A + P_B + \dots = P_{tot}$.

Partial Pressure of Ideal Gases: $P_A = \frac{n_A RT}{V}$ and $P_{tot} = \frac{n_{tot} RT}{V}$ for fixed T, V . Similarly, $V_A = \frac{n_A RT}{P}$ and $V_{tot} = \frac{n_{tot} RT}{P}$ for fixed T, P .

Wet Gasses: Gasses which contain water vapour, to determine specific properties of the gas, use the following table for pressure of water vapour and deduce the relevant information.

$t, ^\circ C$	P, kPa
15.0	1.706
17.0	1.938
19.0	2.198
21.0	2.448
23.0	2.811
25.0	3.170
30.0	4.247
50.0	12.35

Average Speed of Gas Molecules: The molecules in a gas have average speed $v_{avg} = \sqrt{\frac{8RT}{\pi M}}$ and root-mean-square speed $v_{rms} = \sqrt{\frac{3RT}{M}}$ which is the square root of the average of the squares of speeds. Note this means for fixed T lighter molecules travel faster on average and for fixed mass increasing temperature increase average speed of molecules.

Effusion: Molecules escaping from their container through an opening.

Diffusion: Molecules of one gas mixing amongst molecules of another gas.

Graham's Law: For effusion and diffusion processes, the rate of effusion or diffusion is given by $\frac{(\text{rate})_X}{(\text{rate})_Y} = \sqrt{\frac{M_Y}{M_X}}$

STP: Standard temperature and pressure, $273^\circ K$ and 100 kPa .

Molar Volume: $\bar{V} = \frac{V}{n}$. Hydrogen and helium have a higher molar volume than an ideal gas, whereas all other gasses have lower molar volume.

Compressibility Factor: $Z = \frac{PV}{nRT} = \frac{p\bar{V}}{RT}$. Note for an ideal gas $Z = 1$. As pressure increases, Z deviates more, and so the gas is less ideal. Hydrogen and helium have a higher compressibility factor than an ideal gas, whereas all other gasses have lower factor.

van der Waals Equation of State: Where a, b are molecule dependent parameters, $\left(P + \frac{an^2}{V^2}\right)(V - nb) = nRT$ or $\left(P + \frac{a}{\bar{V}^2}\right)(\bar{V} - b) = RT$. b relates to the volume occupied by molecules and a relates the attractive intermolecular forces. a and b tend to increase with the size of the molecule.

Virial Equation of State: $Z = \frac{P\bar{V}}{RT} = 1 + \frac{B}{\bar{V}} + \frac{C}{\bar{V}^2} + \dots$. It can be taken to arbitrarily high-order terms.

Macroscopic Properties of Liquid: Vapour pressure: pressure of vapour that forms above a liquid in a closed container. Normal boiling point (T_{vap}): temperature at which the vapour pressure of the liquid equals 1 atm. Surface tension (γ): Energy required to increase the surface area of a liquid. Viscosity (η): measure of fluid's resistance to flow; speed of flow through a tube is inversely proportional to viscosity.

Stronger intermolecular attractions imply higher boiling point, greater surface tension, higher viscosity, and lower vapour pressure.

Polarizability (α): Provides a measure of the extent to which a the charge cloud of a molecule can be distorted (polarized) by another molecule. Larger molecules have larger polarizabilities. Note that non-polar molecules can still have an induced dipole moment if brought near an ion for instance.

London Dispersion Force (LDF): Since electrons are in constant motion, it is possible to have a greater number of electrons at one end. An asymmetric distribution of electrons yields an instantaneous and temporary dipole moment μ_{inst} , which can induce a dipole μ_{ind} in a neighbouring electron. Larger molecules have larger LDFs. LDF is the weakest intermolecular force, but is always present.

Hydrogen Bonding Force: Hydrogen bonds are special LDFs where a hydrogen atom is bonded to $N, O,$ or F . The H atom carries significant positive charge and will be strongly attracted to a lone pair on another molecule. This force is stronger than dipole-dipole and LDFs but less strong than chemical bonds.

Appendix

Constants and Conversion Factors:

R	$= 8.3145 \text{ kPa L K}^{-1} \text{ mol}^{-1}$	1 atm	$= 101.325 \text{ kPa}$
	$= 8.3145 \text{ J K}^{-1} \text{ mol}^{-1}$	1 atm	$= 1.01325 \text{ bar}$
	$= 0.0831415 \text{ bar L K}^{-1} \text{ mol}^{-1}$	1 atm	$= 760 \text{ Torr}$
	$= 0.082058 \text{ atm L K}^{-1} \text{ mol}^{-1}$	1 atm	$= 760 \text{ mmHg}$
N_A	$= 6.022 \cdot 10^{23} \text{ mol}^{-1}$	1 kPa L	$= 1 \text{ J} = 1 \text{ Pa m}^3$
m_c	$= 9.109 \cdot 10^{-31} \text{ kg}$	1 atm L	$= 101.325 \text{ J}$
e	$= 1.602 \cdot 10^{-19} \text{ C}$	1 bar L	$= 100 \text{ J}$
R_H	$= 2.179 \cdot 10^{-18} \text{ J}$	1 cal	$= 4.184 \text{ J}$
c	$= 2.998 \cdot 10^8 \text{ m s}^{-1}$	$1 \mu\text{m}$	$= 10^{-6} \text{ m}$
h	$= 6.626 \cdot 10^{-34} \text{ J s}$	1 nm	$= 10^{-9} \text{ m}$
π	$= 3.141592654$	1 pm	$= 10^{-12} \text{ m}$
		1 amu	$= 1.6605 \cdot 10^{-27} \text{ kg}$
		$0^\circ C$	$= 273.15 \text{ K}$
		1 m^3	$= 1000 \text{ L}$
		1 J	$= 6.242 \cdot 10^{18} \text{ eV}$

Key Equations:

Section	
0	$n = \frac{m}{M}$
0	$C = \frac{Y}{T}$
2	$\nu = \frac{Y}{T}$
2	$c = \lambda\nu$
2	$E_{photon} = h\nu$
2	$E_{photon} = w + (KE)_e$
2	$E_{release} = E_{upper} - E_{lower}$
2	$E_n = -\frac{R_h}{n^2}$
2	$\lambda_{dB} = \frac{h}{mv}$
2	$KE = \frac{1}{2}mv^2$
2	$p = mv$
2	$\Delta x \Delta p \geq \frac{h}{4\pi}$
4	$\Delta_r H = \sum_{\text{reactants}} BDE's - \sum_{\text{products}} BDE's$
4	$\mu = \delta r$
5	$B.O. = \frac{1}{2}(n_b - n_a)$
6	$PV = nRT$
6	$x_A = \frac{n_A}{n_{tot}}$
6	$P_A = x_A P_{tot}$
6	$n_{tot} = n_A + n_B + \dots$
6	$1 = x_A + x_B + \dots$
6	$P_{tot} = P_A + P_B + \dots$
6	$P_A = \frac{n_A RT}{V}$
6	$P_{tot} = \frac{n_{tot} RT}{V}$
6	$V_A = \frac{n_A RT}{P}$
6	$V_{tot} = \frac{n_{tot} RT}{P}$
6	$v_{avg} = \sqrt{\frac{8RT}{\pi M}}$
6	$v_{rms} = \sqrt{\frac{3RT}{M}}$
6	$\frac{(\text{rate})_X}{(\text{rate})_Y} = \sqrt{\frac{M_X}{M_Y}}$
6	$nRT = \left(P + \frac{an^2}{V^2}\right)(V - nb)$
6	$\lambda = \frac{h}{p}$
6	$RT = \left(P + \frac{a}{\bar{V}^2}\right)(\bar{V} - b)$
6	$\bar{V} = \frac{V}{n}$
6	$Z = \frac{PV}{nRT} = \frac{p\bar{V}}{RT}$
6	$Z = 1 + \frac{B}{\bar{V}} + \frac{C}{\bar{V}^2} + \dots$