CM1102
Chemistry – The Central Science

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Contents

Part I

Environment

1   Light

Light is a form of energy that consists of oscillating *electric* and *magnetic* fields, perpendicular to each other.

$c = \lambda \nu$		
	value	**unit**
c	speed of light $= 2.9979 \times 10^8$	m/s
λ	wavelength	m
ν	frequency	s^{-1} or Hz

1.1   Wave-particle Duality of Light

Due to the wave nature of light, light waves that are *in phase* undergo **constructive interference**.

Light waves that are *out of phase* undergo **destructive interference** — cancelling out each other.

Due to the particle nature of light, when an object is heated, it emits electromagnetic radiation known as **blackbody radiation**.

The wavelength of light emitted changes as the temperature of the object increases.

$E = h\nu$		
	value	**unit**
E	energy	J
h	Planck's constant $= 6.626 \times 10^{-34}$	J/s
ν	frequency	s^{-1} or Hz

Putting the two equations above together, we can interconvert between *energy*, *frequency*, and *wavelength*.

$E = \frac{hc}{\lambda}$		
	value	**unit**
E	energy	J
h	Planck's constant $= 6.626 \times 10^{-34}$	J/s
c	speed of light $= 2.9979 \times 10^8$	m/s
λ	wavelength	m

1.2   Photoelectric Effect

Electrons can only be ejected from a surface if incoming **photons** transfer sufficient energy to the atoms on the metal surface.

Electron-volt (eV) is the amount of *kinetic energy* gained by a single electron accelerating from rest though an electric potential difference of 1 volt in vacuum.

$1 \text{ eV} = 1.602 \times 10^{-19} \text{ J}$

$h\nu = \Phi + KE_{max}$		
	value	**unit**
h	Planck's constant $= 6.626 \times 10^{-34}$	J/s
ν	frequency	s^{-1} or Hz
Φ	work function of the metal	eV
KE_{max}	maximum kinetic energy of the ejected electron	eV

Note that KE_{max} *cannot be negative*, so its minimum value is 0.

1.3   Solar Spectrum

The **electromagnetic spectrum** comprises of, in increasing energy:

1. radio waves,
2. microwaves,
3. infrared,
4. visible,
5. ultraviolet,
6. X-rays,
7. gamma (γ) rays,

White light disperses into a continuous spectrum in the visible spectrum.

However, sunlight is *not continuous* — there are *dark lines* in the **solar spectrum** due to the composition of the sun and the Earth's atmosphere.

1.3.1   Aside: Spectroscopy

The study of the interaction of electromagnetic radiation *with* matter *to obtain information about the identity and structure of substances*.

Absorption spectroscopy reveals that a colored object reflects light of that color and absorbs light of the *opposite color*.

1.3.2   Emission Spectrums

Energy can excite electrons of atoms from their **ground state** to an **excited state**.

In the excited state, electrons have higher energy, and when electrons are relaxed to their ground state, their energy is released as light.

Therefore, the dark lines in the solar spectrum occur due to the absorption of energy by the atoms in the sun's atmosphere.

1.3.3   Rydberg Formula

Used to calculate the wavelengths of a spectral line in hydrogenic species, i.e. H, He$^+$, Li^{2+}, etc.

It generalizes the Balmer series ($n = 2$) for the visible

region, the Lyman series ($n = 1$) for the UV region, and the Paschen series ($n = 3$) for the IR region.

$\frac{1}{\lambda} = R_H \left(\frac{1}{n_1{}^2} - \frac{1}{n_2{}^2} \right)$		
	value	**unit**
λ	wavelength of light emitted	m
R_H	Rydberg constant $= 1.096776 \times 10^7$	m^{-1}
n_1	final electron state (lower energy) $< n_2$	int
n_2	initial electron state (higher energy)	int

2 Energy Levels

In the Bohr model, electrons are allowed only in certain circular orbits around the nucleus, which give rise to the **principal quantum numbers** of n.

Therefore, the energy levels in atomic structures are *quantized*, and the energy levels associated with each electron orbit has a fixed value, hence the term **quantized energy level**.

$E_n = -\frac{2\pi^2 m_e e^4}{n^2 h^2}$		
	value	**unit**
E_n	energy level	J
m_e	electron mass $= 9.1094 \times 10^{-31}$	kg
e	electron charge $= 1.602 \times 10^{-19}$	C
n	principal quantum number	int
h	Planck's constant $= 6.626 \times 10^{-34}$	J$\cdot$sec

Since the only variable is n, the equation is generalizable:

$E_n \approx -2.18 \times 10^{-18} \left(\frac{1}{n^2} \right)$ J
$E_n \approx -13.6 \left(\frac{1}{n^2} \right)$ eV

The **ground state** of electrons, $n = 1$, has the lowest energy level and the most stable.

The **excited states** have values of $n > 1$.

The energy required to excite an electron from the ground state to an excited state $\Delta E = E_{\text{final}} - E_{\text{initial}}$ can be determined using the E_n formula above.

If $\Delta E > 0$, this is known as an **absorption** (or excitation) process, going from a lower energy state to a higher energy state.

Conversely, if $\Delta E < 0$, this is known as an **emission** (or relaxation) process, going from a higher energy state to a lower energy state, which releases energy.

The **ionization continuum** occurs as $n \to \infty, E_n \to 0$, such that energy levels are no longer quantized.

The **ionization energy** is the minimum energy required to remove one electron from a gaseous atom or molecule.

In the case of hydrogen, $\Delta E = E_\infty - E_1 = +13.6$eV.

However, this only applies to 1-electron systems, as we otherwise need to account for electron-electron repulsion, which requires taking into account the nuclear charge, Z^2:

$E_n \approx -2.18 \times 10^{-18} \left(\frac{Z^2}{n^2} \right)$ J
$E_n \approx -13.6 \left(\frac{Z^2}{n^2} \right)$ eV

3 Atomic Structure

Electrons, and all matter, behave like waves, which can undergo constructive and destructive interference.

$\lambda = \frac{h}{mv}$		
	value	**unit**
λ	De Broglie wavelength	m
h	Planck's constant $= 6.626 \times 10^{-34}$	J/s
m	particle mass	kg
v	particle velocity	m/sec

By **Heisenberg's uncertainty principle**, it is impossible to know the precise position of electrons in an atom.

3.1 Quantum Mechanical Model

A **wave equation** describes the behavior of a specific electron in an atom.

The solution to the wave equation is called a **wave function** (ψ), or an **orbital** — the space boundary where there is a 95% chance of finding the electron.

$|\psi|^2$ is proportional to the *probability density* of an electron being found in 3D space.

The position of an electron could be described using *Cartesian coordinates*, using the wave equation $\psi(x, y, z)$.

Alternatively, using *spherical coordinates*, $\psi(r, \theta, \phi)$:

| $\psi(r, \theta, \phi) = \mathbf{R}(r) \times \mathbf{Y}(\theta, \phi)$		
	value	**unit**
$\mathbf{R}(r)$	**radial wave function**	–
r	distance of the electron from the nucleus	–
$\mathbf{Y}(\theta, \phi)$	**angular wave function**	–
θ	vertical angle perpendicular to the x-y plane	–
ϕ	angle on the x-y plane	–

Radial nodes are 3D *spherical* regions where where no electrons are found. The *radial* component of the wave function passes through 0.

Angular nodes are 2D *planar* or 3D *conical* regions where where no electrons are found. The *angular* component of the wave function passes through 0.

In orbital diagrams, the regions corresponding to positive values of the wave function are shaded, while the regions corresponding to negative values are not.

3.1.1 Radial Distribution Function (RDF)

$RDF : 4\pi r^2 \cdot \mathbf{R}(r)^2$		
	value	**unit**
$4\pi r^2$	surface area of a sphere	–
$\mathbf{R}(r)$	radial wave function	–
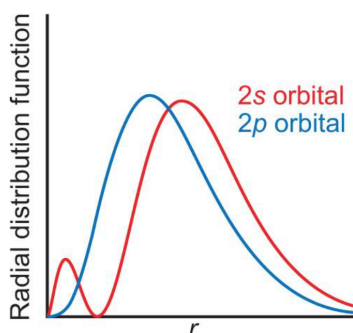
The maximum value of the RDF (i.e. the peak) is the most probable distance of the electron from the nucleus.

Even though the larger peak of the $2p$ orbital is closer than that of the $2s$ orbital, the $2s$ orbital has a higher energy level than the $2p$ orbital due to the **penetration** from the smaller peak.

Therefore, within shells of the same quantum number, orbital energy increases as follows: $s < p < d < f$.

Radial nodes can be counted from the x-intercepts, so from the diagram above, the $2s$ orbital has 1 radial node, while the $2p$ orbital has none.

3.1.2 Quantum Numbers

nl_{m_l}, e.g. $2p_z$		
	value	**range**
n	principal QN	$1, 2, 3, \ldots$
l	magnetic QN	$0, 1, \ldots, n-1$
m_l	orbital angular momentum QN	$-l, \ldots, 0, \ldots, +l$

The principal QN relates to the *size* and *energy* of the orbital — higher principal QNs have larger orbitals with higher energy levels.

The magnetic QN relates to the *shape* of the orbital.

The orbital angular momentum QN relates to the *orientation* of the orbital.

number of angular nodes	l
number of radial nodes	$n-l-1$

3.2 Electronic Configuration

Pauli's exclusion principle: No two electrons in a given atom can have the exact same set of quantum numbers (n, l, m_l, m_s).

Aufbau principle: Electrons are always placed in the lowest energy sublevel available.

Hund's rule: Degenerate orbitals are always filled with single electrons before any of them are doubly occupied, and electrons fill separate orbitals with their spins aligned.

3.3 Trends in Atomic Properties

trend	$\rightarrow$ **period**	$\downarrow$ **group**
atomic radius	decreases	increases
ionization energy	increases	decreases
electron affinity	increases	decreases

Atomic radius is affected by the **effective nuclear charge** (Z_{eff}) of the atom.

$Z_{\text{eff}} = Z - S$		
	value	**unit**
Z	nuclear charge (atomic number)	int
S	shielding constant	real

Ionization energy is the energy required to remove an electron from a gaseous atom or ion.

Electron affinity is the energy released when an electron is added to a gaseous atom to form an anion.