

Exam 1 material begins:

8/28 lecture notes

Chemistry = studying properties of matter by particles that make them up

-91 natural atoms

Structure of atom

- nucleus has lots of dense mass with all + charge (protons and neutrons)
- most of atom is empty space
- number protons of nucleus = number of electrons in orbit
- protons and electrons have equal but opposite charge
- protons and neutrons have similar mass but electrons are much less massive

Elements

- atomic number = number of protons in nucleus
- SYMBOL NO NAME ON PERIODIC TABLES ON TESTS

Isotopes

- different number of neutrons = different isotope
- mass number = # of protons plus number of neutrons

Ions

- has missing or extra electrons
- negative charged ions = anion (extra electrons)
- positively charged ions = cations (missing electrons)
- can only make ions by changing electrons CHANGING PROTONS CHANGES THE ELEMENT

Atomic mass

- weighted value based on natural abundance of isotopes which naturally occur
- unit is amu or grams/mole (g/mol)

Dimensional analysis

- ALWAYS REPORT UNITS IN ANSWERS ON HW AND EXAMS
- unit conversion: $\text{current unit} = \text{desired unit} / \text{current unit}$
- make the fraction equal to one (numerator must = denominator in different units) and cross of both sides
- 1ml = 1cm³
- 1in = 2.54cm

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- 10⁹ nm in meter
- Avogadro's number: a constant which tells exactly how many molecules per mole (1 mole = 6.022×10^{23} particles)
- this number comes from the exact number of atoms in 12 grams of pure carbon-12

- we didn't get avogadro's number in class problem since we assumed no space between molecules which isn't practical
- what we know about atoms: dense nucleus (protons and neutrons) which is positively charged, electrons = protons = atomic number

Goal of chapter 7: determine electrons structure in atom

- what are electrons?
- where are electrons?

Quantum Theory

- classical mechanics = macroscopic level
- quantum theory = microscopic level

Wave nature of light

- light is electromagnetic radiation
- moves in oscillating waves
- propagates through space at 3.00×10^8 m/s
- amplitude of wave determines brightness of light
- wavelength determines color (ROYGBIV)
- violet has smallest wavelength at 400nm
- red has largest wavelength at 750nm
- frequency (cycles/second or s^{-1} or Hz) can also be used to describe color
- frequency = number of crests per given time
- frequency and wavelength are inversely proportional (as wavelength increases frequency decreases)
- frequency (ν) = speed (C)/wavelength (λ)
- HW calculate frequency of violet light
- KNOW ORDER OF ELECTROMAGNETIC SPECTRUM

Interference

- constructive interference = waves in phase and create bigger wave (more amplitude)
- destructive interference = waves cancel each other out and make small or no wave

Diffraction

- when traveling waves hit a hole in a barrier which is similar in size to the wave it diffracts
- if a wave goes through multiple holes the light can interfere with each other
- light behaves as wave in this case with holes but not always in every case

Photoelectric effect

- challenges wave nature of light
- light hits metal and "bounces off" as electrons
- explanation: energy is transferred from light to electrons in metal until it builds up to be larger than the binding energy holding down the electrons and sends it off of metal and excess energy becomes kinetic energy with the electron

- the more transferred energy from light the more kinetic energy with electron
- classical electromagnetic theory would say light is a wave so the energy of the wave is proportional to the amplitude only (predicting the greater intensity of light the more kinetic energy in electrons)
- this was experimented and no relation was found between kinetic energy of electrons but there was a relationship between intensity and number of electrons
- they did find that after a point (ν_0 or threshold frequency) there is a constant increase of kinetic energy as the frequency is increased
- after threshold frequency number of electrons stayed constant when frequency increased
- Einstein concluded that light is acting as a particle in this situation and wave in other cases
- frequency = speed/ wavelength
- energy of particles of light proportional to frequency
- number of particles of light proportional to intensity (more photons of light = more electrons released)

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Einstein: Light can be described as a collection of packets of energy (photons)

- energy of one photon = plancks constant times frequency ($E_{\text{photon}} = h\nu$) → since we concluded energy of light particles proportional to frequency
- energy in = energy out
- in photoelectric problem if there is leftover energy after the binding energy is used it turns into Kinetic energy
- if there is just enough for binding energy then the frequency is the threshold frequency
- since electrons are ejected with kinetic energy → binding energy is usually $< E_{\text{photon}}$
- if binding energy $> E_{\text{photon}}$ then no electrons are released with KE_e
- if binding energy = E_{photon} then electrons are released with no kinetic energy

Can electrons behave as waves as well?

- Atomic spectrocity: exciting elements to cause different colored light (neon signs)
- if electrons obeyed classical physics there would be a full continuous spectrum of light but there are major gaps on the spectrum (it jumps from one color to another with no transition)

The Bohr Model

- energy of atom is quantized
- electrons are at fixed distances from the nucleus
- this explains the spectrum because electrons don't exist between the levels meaning there is no transition between colors
- the further from nucleus the higher energy
- electrons want to lose energy and move closer to nucleus
- energy released in light is energy released when electrons change energy levels

Double slit experiment with electrons

- with one slit electrons show interference pattern of a particle
- with two slits electrons show interference pattern of a wave

-when observed with two slits the electrons went back to acting as a particle

De Broglie relation

- all matter has wavelength but its not relevant unless the matter is tiny
- wavelength is inversely proportional to momentum

Heisenberg's uncertainty principal

-if we observe either position or momentum of electron we will not be certain of the other

Indeterminacy

-unlike in classical mechanics when under the exact same circumstances matter does the same thing in quantum we can only statistically guess what particles will do

Quantum and the atom

-Schrodinger equation: $\hbar\psi = E\psi$

Quantum numbers

- principle quantum number = n (determines size and energy of orbital)
- each orbital has a n value (energy level of electron) which can be any positive integer (1, 2, 3...)
- the smaller the n value the closer to the nucleus (lower energy level)
- angular momentum = L which can be any number from $0 \rightarrow (n-1)$
- when $L = 0 \rightarrow S$ orbital
- when $L = 1 \rightarrow P$ orbital
- when $L = 2 \rightarrow D$ orbital
- when $L = 3 \rightarrow F$ orbital
- magnetic quantum number = m_L which can range from $-L \rightarrow L$
- m_L describes how many of each type of orbital there are [ex. If $L=1$ there are 3 P orbitals since there are 3 m_L values (-1, 0, 1)]
- spin number = m_s describes spin of electron (can only be $\frac{1}{2}$ or $-\frac{1}{2}$)
- since each orbital can only hold 2 electrons if two electrons are in the same orbital they have opposite spin numbers
- when naming orbitals the n value goes first (ex 3s orbital)
- wave function $\psi_n(x) = A \sin(n\pi x/L)$ where $n = 1, 2, 3$, etc.

-helpful video I found: <https://www.youtube.com/watch?v=Aoi4j8es4gQ>

9/6 Chapter 8 reading notes

Periodic table

- atomic mass increase left to right
- similar properties in columns
- periodic properties: atomic radius, ionization energy, electron affinity, density, metallic character
- periodic properties are predictable based on location on periodic table

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Putting quantum numbers together to construct atomic orbitals

- principle level (shell) given by n
- sublevels given by L
- number of orbitals in a sublevel given by m_L

Atomic spectroscopy (take 2)

- light comes from photon with same wavelength as an electron transitioning between two quantum mechanical orbitals
- principle energy levels are closer together the further they are from the nucleus
- if an electron moves up between principal levels it is called absorption or excitation
- when an electron moves down between principle levels it's called emission or relaxation and that's when a photon is released
- Rydberg equation: $\Delta E = E_{\text{final}} - E_{\text{initial}}$
- $-\Delta E_{\text{electron}} = \Delta E_{\text{atom}}$ (they are interchangeable since the electron is the only one energy change at a time)
- absorption: $E_{\text{photon}} = \Delta E_{\text{electron}}$
- $-\Delta E_{\text{electron}} = +E_{\text{photon}}$ (+ E_{photon} : absorption, - E_{photon} : emission)
- Rydberg equation only works for hydrogen**

non-hydrogen atoms

- $E_{\text{absorbed}} = \Delta E_{\text{atom}} = E_{\text{emitted}}$ (conservation of energy)

-energy and wavelength are INVERSELY proportionate

Radial distribution function

- $-\psi^2$ when plotted shows that electrons are likely found in nucleus which is false
- we fix that with radial distribution by starting the probability after the sphere which is the nucleus
- nodes in radial function are a place where there is 0% chance of finding an electron there
- point of origin on graph is not a node since it's the nucleus
- nodes in an orbital = $n-1$

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- all electrons in highest n level are valence electrons

Transition elements

- if it's a transition element and the d sublevel is full those electrons are not valence
- transition metals (d block)
- inner transition metals (f block)
- show different trends in orbital filling
- because of sublevel splitting $4s$ sublevel is lower energy than $3d$ so it fills first
- energy difference is small