Lecture 1: Early chemistry led to the discovery (and classification) of the elements, Dalton’s atomism, and much of which was done by lighting stuff on fire (and balancing combustion reactions).
Lecture 2: Mendeleev organized atoms with periodicity, giving rise to our current periodic table. Avogadro introduced the concept of the mole, and the discovery of the electron brought atomic theory to a new level.
Lecture 3: electrons were not only discovered, but also weighed, generated by a cathode ray tube, and shown to be far away from the center of the atom (planetary model).
Lecture 4: Using Planck and Einstein’s work on quantization, Bohr created a model for the atom that explained how electrons could absorb or emit energy by transitioning between orbits.
Lecture 5: Wave-particle duality: everything is both and quantum mechanics rules the day, with the Schrödinger equation governing the wave function.

Electron Spectroscopy

Photoemission Spectroscopy
Lecture 6: Solving the Schrödinger equation for hydrogen gives us the principle ($n$), azimuthal ($l$), and magnetic ($m$) quantum numbers, with electron spin (up/down) shown separately as a fourth quantum number.

We had the knowledge needed to fill these levels and understand the periodic table.
Lecture 7: Hund’s rule has exceptions but gives us the periodic table with electrons filling their quantum mechanical states; concepts like ionization energy and electron affinity, and trends in the periodic table become understood.
Lecture 8: The properties of ionic solids can be understood by examining electronegativity and electron transfer among atoms, and using Hess’s law and the Born-Haber cycle lattice energies can be computed.
Lecture 9: Lewis dot diagrams provide a simple way to convey how valence electrons bond between atoms, using the octet rule to determine bonding configurations, and electronegativity to think about covalent vs. ionic vs. partially ionic bonds types.
Lecture 10: Deeper way to describe bonding: use quantum mechanics. Valence bond theory involves formation of molecular orbitals using a linear combination of atomic orbitals, to give us bonding, non-bonding, and anti-bonding states.
Lecture 11: Hybridization between atomic orbitals allows some elements to lower their energy when they bond to other atoms by creating new types of bonding orbitals, and VSEPR is a formalism to predict molecular shapes.

\( \text{(b) } \text{C}_2\text{H}_4 \pi \text{ bonding} \)

\( \text{(b) } \text{C}_2\text{H}_2 \pi \text{ bonding} \)

VSEPR gives us these shapes.
Lecture 12: Beyond ionic and covalent bonds, there are weaker types of bonds — van der Waals, London dispersion, and Hydrogen bonding — that play an enormous role in material properties.

ethanol
B.P. = 78.5°C

methoxymethane
B.P. = -24.8°C
Lecture 13: The nature of metallic bonding and “sea” of electrons explains melting/boiling point trends in metals, the impact of interstitial or substitutional impurities, and the formation of bands in the solid.

(a) Group 1 metal
(b) Group 2 metal

Antibonding MOs

Energy band

Band gap

Valence bands

Core bands

Empty
Filled

3p
3s

2p

2s

1s

Al
Mg
Na
Lecture 14: Semiconductors, unlike metals, have a gap between the valence (filled) and conduction (empty) bands, but electrons can be put into or removed from these bands via doping, enabling all modern electrons
Lecture 15: Crystals are solids with periodically repeating atoms or molecules, and of the 7 possible crystal systems we focus on cubic crystals for which there are 3 Bravais lattices: simple cubic, body-centered cubic, and face-centered cubic.
Lecture 16: The lattice plus the basis fully describe the crystal, and both a point in space as well as a plane can be specified, with Miller indices giving a convenient way to write crystal planes, and always being mindful that crystallographers are picky about notation.
Lecture 17: Röntgen discovered a new type of Ray, the X-Ray, and Moseley used these rays to probe atoms and for the first time give fundamental meaning to atomic number.
Lecture 18: Using the Bragg condition and selection rules, x-ray diffraction can determine the crystal structure of a solid, by measuring the distance between repeating planes.

$n\lambda = 2d \sin \theta$

- Aluminum XRD
- Cu Target $K_α (\lambda=1.54 \text{ Å})$
Lecture 19: X-ray diffraction can use either the sharp (K, L) lines or the continuous Bremsstrahlung radiation, with the latter used by Laue to prove simultaneously the wave-like nature of X-rays and the space-lattice structure of crystals.
Lecture 20: no crystals are perfect; different types of important point defects include vacancies, substitutional, and interstitial; in ionic crystals we have special names, and both Schottky and Frenkel defects must maintain charge balance.
Lecture 21: moving from zero to one-dimensional defects, the movement of line defects (dislocations) in crystals allows atomic planes to slide past one another with low energy, giving metals their malleability; cold-working creates a large number of dislocations that prevent further dislocation movement making the crystal stronger.
Lecture 22: a glass is an amorphous solid material that forms because atoms or molecular units cannot reach their crystal lattice sites (like atoms playing musical chairs); with solidification at the glass transition temperature and properties like density dependent on cooling rate.
Lecture 23: the properties of glass can be engineered by tempering (as in Prince Rupert’s drop), by introducing network modifiers with $O^{2-}$ as the “spaghetti knife” (as in soda lime glass), or by ion exchange (as in Gorilla Glass).
Lecture 24: the rate of a reaction depends on concentrations according to the rate law (measured experimentally), and the rate constant has an Arrhenius dependence on temperature.

**zeroth-order reaction**

Equation 14.15: \( \text{rate} = - \frac{\Delta[A]}{\Delta t} = k \)

**first-order reaction**

Equation 14.16: \([A] = [A]_0 - kt\)

Equation 14.19: \( \text{rate} = - \frac{\Delta[A]}{\Delta t} = k[A] \)

Equation 14.20: \([A] = [A]_0 e^{-kt}\)

Equation 14.21: \(\ln[A] = \ln[A]_0 - kt\)

**second-order reaction**

Equation 14.22: \( \text{rate} = - \frac{\Delta[A]}{\Delta t} = k[A]^2 \)

Equation 14.23: \( \frac{1}{[A]} = \frac{1}{[A]_0} + kt \)
Lecture 25: Diffusion (whether a dye molecule, a vacancy, an interstitial, etc.) depends on a concentration gradient and a diffusion constant according to Fick’s first law for the steady-state case and Fick’s 2nd law for time-dependent diffusion.
Lecture 26: Solubility in aqueous solutions depends on the solute/solvent interactions and for ionic compounds solubility rules have been determined and generally apply.
Lecture 27: Arrhenius acids donate protons and bases donate $\text{OH}^-$, while Bronsted-Lowry acids and bases are proton donors and acceptors; when reacted acids and bases form conjugate pairs; the number of strong acids and bases is low, most are weak.

\[
\begin{align*}
\text{H-Cl} &+ \text{H}_2\text{O} \overset{\rightleftharpoons}{\longrightarrow} \text{Cl}^{(-)} + \text{H}_3\text{O}^{(+)}, \\
\text{(acid)} &\quad \text{(base)} & \quad \text{(base)} &\quad \text{(acid)} \\
\text{H}_3\text{N}: &+ \text{H}_2\text{O} \overset{\rightleftharpoons}{\longrightarrow} \text{NH}_4^{(+)} + \text{HO}^{(-)}, \\
\text{(base)} &\quad \text{(acid)} & \quad \text{(acid)} &\quad \text{(base)}
\end{align*}
\]

Acid + Base $\rightarrow$ Water + Salt

Autoionization of Water
Lecture 28: The solubility product is a form of the more general equilibrium constant, and is given by the product of concentrations of ions in a saturated solution; for pure water the solubility product is $10^{-14}$, and gives us the pH and pOH scales.

Soren Sorensen “the power of Hydrogen”

### pH of Common Substances

<table>
<thead>
<tr>
<th>ACIDIC</th>
<th>NEUTRAL</th>
<th>ALKALINE OR BASIC</th>
</tr>
</thead>
<tbody>
<tr>
<td>0</td>
<td>1</td>
<td>2</td>
</tr>
<tr>
<td>Battery Acid (Hydrochloric)</td>
<td>Stomach Acid</td>
<td>Lemon Juice, Vinegar</td>
</tr>
<tr>
<td>Coke and Orange Juice</td>
<td>Apple, Dr. Pepper Soda</td>
<td>Tomato Juice, Beer</td>
</tr>
<tr>
<td>3</td>
<td>4</td>
<td>5</td>
</tr>
<tr>
<td>Acid Coffee, Pepsi Bismol</td>
<td>Healthy Skin, Hair and Nails</td>
<td>Urine, Saliva, Milk</td>
</tr>
<tr>
<td>6</td>
<td>7</td>
<td>8</td>
</tr>
<tr>
<td>“Pure” Water, Blood Shampoos</td>
<td>Baking Soda, Seawater, Eggs</td>
<td>Perm Solutions (8.5 to 9.5)</td>
</tr>
<tr>
<td>9</td>
<td>10</td>
<td>11</td>
</tr>
<tr>
<td>Toothpaste, Hand Soap</td>
<td>Milk of Magnesia, Mild Detergent</td>
<td>Household Ammonia and Cleaners</td>
</tr>
<tr>
<td>12</td>
<td>13</td>
<td>14</td>
</tr>
<tr>
<td>Bleach, Oven Cleaner</td>
<td>Liquid Drain Cleaner, Caustic Soda</td>
<td></td>
</tr>
</tbody>
</table>

$K_{sp} = [F^-]^2[Ca^{2+}]$

Solubility Product
Lecture 29: In organic chemistry, we have a systematic way to name molecules depending on the type and number of carbon-carbon bonds, and types and positions of additional groups.
Lecture 30: Polymers ("Many Units") are macromolecules that can be synthesized in two important ways: 1) free radical polymerization where a radical is used to initiate a chain reaction, and 2) condensation polymerization where two end groups contain OH and H that form water while linking together.

[Chemical structures and diagrams are shown here, illustrating ethylene and polyethylene, as well as other polymerization reactions.]
Lecture 31: Polymers can be engineered to possess a wide range of optical, electronic, mechanical, and chemical properties, with structure dictating behavior; molecular weight (degree of polymerization), processing conditions (changing density/crystallinity), branching, tacticity, cross-linking, and chemistry are among tuning parameters.
Lecture 32: Bio-Polymers are made by nature via condensation polymerization, and are nearly infinitely flexible due to the large number and sequence of molecular units used, like 20 amino acids in proteins 1000's of units long and 4 base pairs in DNA 3 billion units long.
Lecture 33: Unary phase diagrams show equilibrium phases of a single-component as a function of pressure and temperature, with phase boundaries, critical points, and the triple point all indicated; a microscopic view of the phase transition reveals how the coexistence curve depends on each variable.

Otto and the power of 1 atm
Lecture 34: Binary phase diagrams show phase equilibria as a function of composition and temperature for two components, with the 2-phase regions no longer just a line (as in the unary case) but rather a range of composition and temperature; tie lines and the lever rule tell us about composition and phase fraction inside a 2-phase region.
Lecture 35: When two components are not miscible, a miscibility gap is present in the phase diagram, and for the case of miscible liquids but immiscible solids, this often leads to a eutectic-type phase diagram and a eutectic point (lowest temperature for liquid to form)