Important Dates:

HW 2 due today
HW 3 due next Thursday
Quiz 3 today
EXAM 1 NEXT THURSDAY
Polar Ozone Depletion – The “Ozone Hole”
### Atomic Structure and Periodicity

#### Table 2.3: Isotopes of Hydrogen

<table>
<thead>
<tr>
<th>Isotope</th>
<th>Isotopic Symbol</th>
<th>Number of Protons</th>
<th>Number of Neutrons</th>
<th>Sum of Protons and Neutrons</th>
</tr>
</thead>
<tbody>
<tr>
<td>hydrogen, H-1</td>
<td>$^1\text{H}$</td>
<td>1</td>
<td>0</td>
<td>1</td>
</tr>
<tr>
<td>deuterium, H-2</td>
<td>$^2\text{H}$</td>
<td>1</td>
<td>1</td>
<td>2</td>
</tr>
<tr>
<td>tritium, H-3</td>
<td>$^3\text{H}$</td>
<td>1</td>
<td>2</td>
<td>3</td>
</tr>
</tbody>
</table>
Molecules and Models

\[ S = N - A \]

- Example: hydrogen fluoride, HF
  - Needed (N): hydrogen needs 2e\textsuperscript{-} and fluorine needs 8e\textsuperscript{-}.
    - \( N = 2 + 8 = 10 \)
  - Available (A): hydrogen has 1e\textsuperscript{-} and fluorine has 7e\textsuperscript{-}.
    - \( A = 1 + 7 = 8 \)
  - Shared (S) = N – A = 10 – 8 = 2
  - So 2e\textsuperscript{-} shared between H and F, a single covalent bond
Molecules and Models

\[ \text{H} \cdot \text{F} \]

- Lone Pairs
- Shared Electrons

\[ \text{H} - \text{F} \]

- Lone Pairs are “understood”
- Single Covalent Bond
Molecules and Models

- Resonance Forms
  - Structures that are hypothetical extremes of electron arrangements that do not exist exactly as represented by any one Lewis structure.

- Ozone’s structure is ‘in-between’ the two resonance structures.
“Free Radicals”

Species with an odd number of electrons – or “unpaired” electrons – are referred to as Free Radicals, or Radicals.

Radicals are extremely reactive, because that single unpaired electron will do just about anything to make a pair.

Often (but not always), we indicate that a species is a radical by representing the unpaired electron with a single dot: •

NO•, NO₂•, Cl•, Br•, OH•, ...
Waves of Light

- Longer wavelength = Lower frequency
- Shorter wavelength = Higher frequency

$\lambda$, wavelength
Waves of Light
Waves of Light

Visible Light

wavelength (\(\lambda\)) in nanometers

400 450 500 550 600 650 700
Waves of Light

\[ \nu = \frac{c}{\lambda} \quad \text{and} \quad c = 3.00 \times 10^8 \frac{m}{s} \]

green light has \( \lambda = 550 \text{ nm} \)

\( 10^9 \text{ nm} = 1 \text{ m} \) or \( 1 \text{ nm} = 10^{-9} \text{ m} \)

\( \lambda = 550 \text{ nm} = 550 \times 10^{-9} \text{ m} = 5.50 \times 10^{-7} \text{ m} \)

\[ \nu = \frac{c}{\lambda} = \frac{3.00 \times 10^8 \frac{m}{s}}{5.50 \times 10^{-7} \text{ m}} = 5.45 \times 10^{14} / \text{s} = 5.45 \times 10^{14} \text{ s}^{-1} \]
Waves of Light

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Waves of Light
Waves of Light – Particles of Light?

• In the early 1900s, physical science was turned upside down by the development of Quantum Mechanics (as opposed to Classical Mechanics)
• Quantum Mechanics is required in order to apply physics to the world on the scale of atoms and molecules
• One of the most shocking postulates of Quantum Mechanics was that light simultaneously possesses properties of a wave and of a particle
  – “Wave/particle duality”
• A “particle” of light is called a photon, and while it definitely possesses the frequency and wavelength of waves, it also carries a specific amount of energy per photon
Energy of Photons

\[ E = h\nu = \frac{hc}{\lambda} \]

\[ h = 6.63 \times 10^{-34} \text{ J} \cdot \text{s} \]

For green light with \( \nu = 5.45 \times 10^{14} \text{ s}^{-1} \)

\[ E = \left( 6.63 \times 10^{-34} \text{ J} \cdot \text{s} \right) \left( 5.45 \times 10^{14} \text{ s}^{-1} \right) \]

\[ E = 3.61 \times 10^{-19} \text{ J} \]
The Interaction of Light with Molecules

• It turns out that the energy in one photon of ultraviolet light has approximately the same energy as a molecular bond!

• If a molecule is struck by a photon of the right energy – that is, of the right frequency! – the molecular bonds will break
The Interaction of UV Light with DNA

http://earthobservatory.nasa.gov/Library/UVB
Ultraviolet Radiation From The Sun

The diagram illustrates the energy intensity of ultraviolet radiation from the Sun at different wavelengths. The X-axis represents the wavelength (nm) ranging from 280 to 360 nm, and the Y-axis represents the energy intensity (J/m².s) ranging from $10^{-6}$ to $10^{0}$.

- **Above the atmosphere**: The intensity remains relatively constant due to the Earth's atmosphere blocking most of the ultraviolet radiation.
- **At surface of the Earth**: The intensity is significantly lower due to the absorption of ultraviolet B (UV-B) radiation by ozone ($O_3$) in this region.

UV-A radiation (blue line) is absorbed more by ozone compared to UV-B radiation (red line) above the atmosphere, with a sharp increase in absorption at wavelengths around 300 nm.
Biological Effects of Ultraviolet Radiation

Melanin production
Sunburn
Premature skin aging
Skin cancer

Absorption by O₃ in this region

Wavelength (nm)
Biological Effects of Ultraviolet Radiation
Scientists have determined that a given % reduction in O$_3$ concentration will produce twice that % increase in skin cancer.
Polar Ozone Depletion – The “Ozone Hole”
Polar Ozone Depletion – The “Ozone Hole”
Sept. 2001

Sept. 2002

Sept. 2003
Ozone: What and Where Is It?

- Ozone, $\text{O}_3$, an allotrope of oxygen

$\text{Energy} + 3 \text{O}_2 \rightarrow 2 \text{O}_3$
Ozone: What and Where Is It?

- Ozone layer
- Stratosphere
- Ozone concentration
- Troposphere
- Altitude (km)
- Altitude (miles)
- Ozone increases from pollution
Ozone: What and Where Is It?

Table 2.4  Categories and Characteristics of UV Radiation

<table>
<thead>
<tr>
<th>Radiation</th>
<th>Wavelength Range</th>
<th>Relative Energy</th>
<th>Comments</th>
</tr>
</thead>
<tbody>
<tr>
<td>UV-A</td>
<td>320–400 nm</td>
<td>Least energetic of these three UV categories</td>
<td>Least damaging, reaches Earth’s surface in greatest amount</td>
</tr>
<tr>
<td>UV-B</td>
<td>280–320 nm</td>
<td>More energetic than UV-A, less energetic than UV-C</td>
<td>More damaging than UV-A, less damaging than UV-C, most absorbed by ozone in the stratosphere</td>
</tr>
<tr>
<td>UV-C</td>
<td>200–280 nm</td>
<td>Most energetic of these three categories</td>
<td>Most damaging of these three, but not a problem because totally absorbed by oxygen and ozone in stratosphere</td>
</tr>
</tbody>
</table>
The Oxygen/Ozone Screen

\[ O_2 \xrightarrow{\lambda \leq 242 \text{ nm}} 2O \]

\[ O_3 \xrightarrow{\lambda \leq 320 \text{ nm}} O_2 + O \]
The Oxygen/Ozone Screen

• These two reactions occurring in the stratosphere effectively filter the most dangerous wavelengths of UV radiation.

• Ozone in the stratosphere is continuously being created and destroyed

• Steady State
  – A condition in which a dynamic system is in balance so that there is no net change in concentration of the major species involved.

• Chapman cycle
The Chapman Cycle

$O_2 \xrightarrow{UV \text{ photons} \ (\lambda \leq 242 \text{ nm})} 2O$  
new $O$ fed into cycle

O + $O_2$ \quad \text{collisions, fast} \quad \{ \text{subcycle} \}

$O_3 + O \xrightarrow{\text{collisions} \ (\lambda \leq 320 \text{ nm})} 2O_2$

(O$_3$ removed from cycle)

Both of these reactions give off heat
Stratospheric Ozone Destruction – A Polar Phenomenon

Antarctic ozone minimum (60°–90° S)

1979–1992 Nimbus 7 TOMS
1979–1994 Meteor 3 TOMS
1995 (no TOMS in orbit)
1996–2003 Earth Probe TOMS
Stratospheric Ozone Destruction –
Also A Global Phenomenon

Trend: 1926–1973
10.1%/decade

22.9%/decade

TOMS ozone

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