Emission Spectrum of Hydrogen
The Periodic Table

Figure 2.5  Dmitri Mendeleev’s 1872 periodic table. The spaces marked with blank lines represent elements that Mendeleev deduced existed but were unknown at the time, so he left places for them in the table. The symbols at the top of the columns (e.g., R^2O and RH^4) are molecular formulas written in the style of the 19th century.
# Naturally Occurring Man-Made

**The Periodic Table**

|   | 1 | 2 | 3 | 4 | 5 | 6 | 7 | 8 | 9 | 10 | 11 | 12 | 13 | 14 | 15 | 16 | 17 | 18 |
| 1 | H 1.0079 | 2 | He 4.0026 |
| 3 | Li 6.941 | 4 | Be 9.0122 |
| 5 | B 10.811 |
| 6 | C 12.011 |
| 7 | N 14.007 |
| 8 | O 15.999 |
| 9 | F 18.998 |
| 10 | Ne 20.180 |
| 11 | Na 22.990 | Mg 24.305 |
| 12 | Al 26.982 |
| 13 | Si 28.086 |
| 14 | P 30.974 |
| 15 | S 32.065 |
| 16 | Cl 35.453 |
| 17 | Ar 39.948 |
| 18 | K 39.102 |
| 19 | Ca 40.078 |
| 20 | Sc 44.956 |
| 21 | Ti 47.867 |
| 22 | V 50.942 |
| 23 | Cr 52.000 |
| 24 | Mn 54.938 |
| 25 | Fe 55.845 |
| 26 | Co 58.933 |
| 27 | Ni 58.693 |
| 28 | Cu 63.546 |
| 29 | Zn 65.409 |
| 30 | Ga 69.723 |
| 31 | Ge 72.64 |
| 32 | As 74.922 |
| 33 | Se 78.96 |
| 34 | Br 79.994 |
| 35 | Kr 83.798 |
| 36 | Rb 85.468 |
| 37 | Sr 87.62 |
| 38 | Y 88.906 |
| 39 | Zr 91.224 |
| 40 | Nb 92.906 |
| 41 | Mo 95.94 |
| 42 | Tc 98.906 |
| 43 | Ru 101.07 |
| 44 | Rh 102.91 |
| 45 | Pd 106.42 |
| 46 | Ag 107.87 |
| 47 | Cd 112.41 |
| 48 | In 114.82 |
| 49 | Sn 118.71 |
| 50 | Sb 121.76 |
| 51 | Te 127.60 |
| 52 | I 126.90 |
| 53 | Xe 131.29 |
| 54 | Cs 132.91 |
| 55 | Ba 137.33 |
| 56 | La 138.91 |
| 57 | Ce 140.12 |
| 58 | Pr 140.91 |
| 59 | Nd 144.24 |
| 60 | Pm 150.36 |
| 61 | Sm 151.96 |
| 62 | Eu 157.25 |
| 63 | Tb 158.93 |
| 64 | Dy 162.50 |
| 65 | Ho 164.93 |
| 66 | Er 167.26 |
| 67 | Tm 168.93 |
| 68 | Yb 173.04 |
| 69 | Lu 174.97 |

Lanthanide series:
- 57 La (138.91)
- 58 Ce (140.12)
- 59 Pr (140.91)
- 60 Nd (144.24)
- 61 Pm (150.36)
- 62 Sm (151.96)
- 63 Eu (157.25)
- 64 Gd (158.93)
- 65 Tb (162.50)
- 66 Dy (164.93)
- 67 Ho (167.26)
- 68 Er (168.93)
- 69 Tm (173.04)
- 70 Yb (174.97)

Actinide series:
- 89 Ac (227)
- 90 Th (232.04)
- 91 Pa (231.04)
- 92 U (238.03)
- 93 Np (237)
- 94 Pu (244)
- 95 Am (243)
- 96 Cm (247)
- 97 Bk (251)
- 98 Cf (252)
- 99 Es (257)
- 100 Fm (258)
- 101 Md (258)
- 102 No (259)
- 103 Lr (262)
The electron in a hydrogen atom travels around the nucleus in a circular orbit.

The energy of the electron in an orbit is proportional to its distance from the nucleus. The further the electron is from the nucleus, the more energy it has.

Only a limited number of orbits with certain energies are allowed. In other words, the orbits are quantized.

The only orbits that are allowed are those for which the angular momentum of the electron is an integral multiple of Planck's constant divided by $2\pi$.

Any object moving along a straight line has a momentum equal to the product of its mass ($m$) times the velocity ($v$) with which it moves. An object moving in a circular orbit has an angular momentum equal to its mass ($m$) times the velocity ($v$) times the radius of the orbit ($r$). Bohr assumed that the angular momentum of the electron can take on only certain values, equal to an integer times Planck's constant divided by $2\pi$.

$$mv = n \left( \frac{h}{2\pi} \right)$$

$$n = 1, 2, 3...$$

$$\Delta E = R_H \left[ \frac{1}{n_1^2} - \frac{1}{n_2^2} \right]$$
Electromagnetic Radiation has associated with it only discrete energies (quantized) ie. light is an electromagnetic wave

\[ E = h\nu \]

Schrödinger wave equation

\[ H\Psi = E\Psi \]

Electromagnetic Radiation can exhibit particle like behaviour

Wave-Particle Duality proposed by De Broglie

\[ \lambda = \frac{h}{mv} \]

Complex physics and mathematics! BUT chemists want to know what it says about molecules not maths
Shapes of orbitals

Why do they take this shape? Quantum Mechanics!
For a hydrogen atom the energies are ordered purely by quantum numbers.

So the 1s orbital is the lowest in energy.

For n = 2 all orbitals (2s and 2p) are the same in energy and said to be *degenerate*.

$n = \infty$ is the ionisation energy i.e. the energy required to remove an electron.

The scale shows a more negative energy as we go to lower quantum numbers – more stable.

For hydrogen the electron is accommodated in the lowest energy orbital. This is known as its *ground state*.

The ground state electronic structure of hydrogen is \(1s^1\).

An electron can be raised in energy (promoted) to an orbital of higher energy. *This is an excited state.*
Energies of orbitals

Emission Spectrum of Hydrogen
Flame tests of other elements

Na$^+$  K$^+$  Li$^+$  Ba$^{2+}$
Filling of electrons

An aid to remember the order
Electronic structure of periods

Valence orbitals – those electrons that participate in chemistry – the highest energy electrons

Core orbitals – those that do not participate in the chemistry – held tightly to the nucleus

The octet rule can now be understood!
The Octet Rule: Atoms try to obtain the noble gas configuration by loss or gain of electrons. How does that work?

Lithium has 3 electrons: \(1s^22s^1\)

If it loses an electron to form Li\(^+\): \(1s^2\)

\(\Rightarrow \text{Li}^+ \equiv \text{He}\)
The Octet Rule:
Atoms try to obtain the noble gas configuration by loss or gain of electrons.

Fluorine has 7 electrons: $1s^22s^22p^5$

If it gains an electron to form $F^-$: $1s^22s^22p^6$

$\Rightarrow F^- \equiv Ne$
The Octet Rule: Atoms try to obtain the noble gas configuration by loss or gain of electrons.

Carbon has 6 electrons: $1s^22s^22p^2$

Does it gain 4 electron or lose 4 electrons?

Answer is neither – it shares its electrons.
Electronic Configurations

The Octet Rule:
Atoms try to obtain the noble gas configuration by loss or gain of electrons.

Carbon has 6 electrons: $1s^22s^22p^2$

Carbon in CH$_4$ shares 4 electrons with hydrogen so as to obtain its octet.
The Periodic Table

<table>
<thead>
<tr>
<th>Period</th>
<th>Trend in Properties</th>
<th>Row</th>
<th>Similar Properties</th>
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- **s block**
- **p block**
- **d block**
- **f block**

**Notable Elements**
- **Lanthanide series**: lanthanum (La), actinium (Ac), etc.
- **Actinide series**: lawrencium (Lr), etc.

**Periodic Table Grid**

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</tbody>
</table>

**Notable Elements**
- **He**: Helium
- **Li**: Lithium
- **Be**: Beryllium
- **B**: Boron
- **C**: Carbon
- **N**: Nitrogen
- **O**: Oxygen
- **F**: Fluorine
- **Ne**: Neon
- **Na**: Sodium
- **Mg**: Magnesium
- **Al**: Aluminum
- **Si**: Silicon
- **P**: Phosphorus
- **S**: Sulfur
- **Cl**: Chlorine
- **Ar**: Argon
- **K**: Potassium
- **Ca**: Calcium
- **Sc**: Scandium
- **Ti**: Titanium
- **V**: Vanadium
- **Cr**: Chromium
- **Mn**: Manganese
- **Fe**: Iron
- **Co**: Cobalt
- **Ni**: Nickel
- **Cu**: Copper
- **Zn**: Zinc
- **Ga**: Gallium
- **Ge**: Germanium
- **As**: Arsenic
- **Se**: Selenium
- **Br**: Bromine
- **Kr**: Krypton
- **Rb**: Rubidium
- **Sr**: Strontium
- **Y**: Yttrium
- **Zr**: Zirconium
- **Nb**: Niobium
- **Mo**: Molybdenum
- **Tc**: Technetium
- **Ru**: Ruthenium
- **Rh**: Rhodium
- **Pd**: Palladium
- **Ag**: Silver
- **Cd**: Cadmium
- **In**: Indium
- **Sn**: Tin
- **Sb**: Antimony
- **Te**: Tellurium
- **I**: Iodine
- **Xe**: Xenon
- **Cs**: Cesium
- **Ba**: Barium
- **La**: Lanthanum
- **Ac**: Actinium
- **Th**: Thorium
- **Pa**: Protactinium
- **U**: Uranium
- **Np**: Neptunium
- **Pu**: Plutonium
- **Am**: Americium
- **Cm**: Curium
- **Bk**: Berkelium
- **Cf**: Californium
- **Es**: Einsteinium
- **Fm**: Flerovium
- **Md**: Moscovium
- **No**: Nihonium
- **Lr**: Lawrencium

**Notable Characteristics**
- Periodic Table organized by atomic number, electron configuration, and chemical properties.
- Trends in properties such as atomic radius, ionization energy, electronegativity, and metallic character across periods and groups.
- Lanthanide and Actinide series highlighting similarities in properties due to their partially filled electron shells.

**Diagram Details**
- The diagram highlights the periodic table with a focus on the s, p, d, and f blocks.
- Key elements are circled to emphasize their position and significance in the periodic system.
LiCO₃ pills for mood disorders

Potassium is found in foods

Cs (and Rb) used in clocks

Li

Na (145g)

K

Rb and Cs
The energy required to completely remove an electron from an atom in the gas phase:

$$E_{(g)} \rightarrow E^+_{(g)} + e^-$$

First ionisation enthalpies (kJ mol$^{-1}$) for the elements Hydrogen to Potassium.
Can we use ionisation energies to rationalise Na$^+$, Mg$^{2+}$, Al$^{3+}$?

Successive ionisation energies
Electronegativity is defined as the power of an atom in a molecule to attract electrons to itself. This is a very powerful principle for understanding the nature of the elements and the types of compounds they form with each other. Pauling assigned the most electronegative element, F, to 4. He noticed that the bond energy $E(AB)$ in a molecule $AB$ is always greater than the mean of the bond energies $E(AA) + E(BB)$ in the homonuclear species $AA$ and $BB$. His argument was that in an "ideal" covalent bond $E(AB)$ should equal this mean, and that the "excess" bond energy is caused by electrostatic attraction between the partially charged atoms in the heteronuclear species $AB$. 
Electronegativity

The 3rd dimension of the periodic table?
Periodic Trends: As you go across a period the electronegativity increases. As you go down a group, electronegativity decreases.

Explaining the Trends in Electronegativity

The attraction that a bonding pair of electrons feels for a particular nucleus depends on:

• the number of protons in the nucleus
• the distance from the nucleus
• the number (and type) of inner electrons.
<table>
<thead>
<tr>
<th>Oxoanion</th>
<th>Name</th>
<th>Oxoacid</th>
<th>Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>ClO(^{-})</td>
<td>hypochlorite</td>
<td>HClO(_{2})(aq)</td>
<td>hypochlorous acid</td>
</tr>
<tr>
<td>ClO(_{2})(^{-})</td>
<td>chlorite</td>
<td>HClO(_{3})(aq)</td>
<td>chlorous acid</td>
</tr>
<tr>
<td>ClO(_{3})(^{-})</td>
<td>chlorate</td>
<td>HClO(_{4})(aq)</td>
<td>chloric acid</td>
</tr>
<tr>
<td>ClO(_{4})(^{-})</td>
<td>perchlorate</td>
<td>HClO(_{4})(aq)</td>
<td>perchloric acid</td>
</tr>
<tr>
<td>NO(_{2})(^{-})</td>
<td>nitrite</td>
<td>HNO(_{2})(aq)</td>
<td>nitrous acid</td>
</tr>
<tr>
<td>NO(_{3})(^{-})</td>
<td>nitrate</td>
<td>HNO(_{3})(aq)</td>
<td>nitric acid</td>
</tr>
<tr>
<td>SO(_{3})(^{2-})</td>
<td>sulfite</td>
<td>H(<em>{2})SO(</em>{3})(aq)</td>
<td>sulfurous acid</td>
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<tr>
<td>SO(_{4})(^{2-})</td>
<td>sulfate</td>
<td>H(<em>{2})SO(</em>{4})(aq)</td>
<td>sulfuric acid</td>
</tr>
<tr>
<td>HSO(_{3})(^{-})</td>
<td>hydrogen sulfite</td>
<td>H(<em>{2})SO(</em>{3})(aq)</td>
<td>sulfurous acid</td>
</tr>
<tr>
<td>HSO(_{4})(^{-})</td>
<td>hydrogen sulfate</td>
<td>H(<em>{2})SO(</em>{4})(aq)</td>
<td>sulfuric acid</td>
</tr>
</tbody>
</table>

if oxoanion ends in "ite" acid ends in "ous"

if oxoanion ends in "ate" acid ends in "ic"
Bonding

Haemoglobin

\[
\text{CH}_3\text{N} - \text{CH} = \text{CH}_2
\]

Chlorophyll

\[
\text{Mg} \quad \text{CH}_3\text{C} = \text{O}
\]

\(d\)-methamphetamine

\[
\begin{array}{c}
\text{C} \quad \text{H} \\
\text{OCH}_3
\end{array}
\]

\(l\)-methamphetamine

\[
\begin{array}{c}
\text{C} \quad \text{H} \\
\text{H}_3\text{C}
\end{array}
\]
How can we use the ideas previously discussed to understand bonding?

Valence orbitals – those electrons that participate in chemistry – the highest energy electrons.

Core orbitals – those that do not participate in the chemistry – held tightly to the nucleus

e.g. F $1s^22s^22p^5$
two major types of bond

1. COVALENT
2. IONIC

Covalent is a sharing of electrons to form a bond

Ionic is loss/gain of electrons
We can use the difference in electronegativity to understand covalent and ionic bonding.
Lewis Structures

What is a bond?
- Sharing of electrons
- Covalent bond, bonding electrons localised, or fixed, between two atoms

Electrons that are not shared are localised as lone pairs

Lewis theory states that all atoms are trying to achieve a noble gas configuration \( \Rightarrow \) OCTET rule

Some rules for Lewis dot diagrams:

Only use valence electrons

Under most circumstances symmetrical geometry is correct!

Oxygen is commonly and Hydrogen always peripheral

Arrange electrons so that all non-H atoms obtain an octet (exceptions for elements in the 3\textsuperscript{rd} and 4\textsuperscript{th} row)
1 – Determine the total number of valence electrons
   Neutral complexes sum the valence electrons
   Cationic complexes *subtract* the charge
   Anionic complexes *add* the charge

2 – Draw the skeletal structure with single bonds. (H is NEVER a central atom)

3 – Place pairs of electrons around the outermost atom

4 – Place any surplus electrons on the central atom

5 – If the central atom does NOT have 8 electrons form a double bond
What are the structures of ionic solids e.g. NaCl?

- Can be thought of as effectively packed arrays of ions

- Efficient means maximising the contacts with oppositely charged ions

The structure of Sodium Chloride shows a coordination number of 6.

The structure of Cesium Chloride shows a coordination number of 8.