8.1 Lewis Dot Symbols

• Valence electrons determine an element’s chemistry.

• *Lewis dot symbols* represent the valence electrons of an atom as dots arranged around the atomic symbol.

• Most useful for main-group elements
Lewis Dot Symbols of the Main Group Elements

<table>
<thead>
<tr>
<th>1A</th>
<th>2A</th>
<th>3A</th>
<th>4A</th>
<th>5A</th>
<th>6A</th>
<th>7A</th>
<th>8A</th>
</tr>
</thead>
<tbody>
<tr>
<td>•H</td>
<td>•Be</td>
<td>•B</td>
<td>•C</td>
<td>•N</td>
<td>•O</td>
<td>•F</td>
<td>•Ne</td>
</tr>
<tr>
<td>•Li</td>
<td>•Mg</td>
<td>•Al</td>
<td>•Si</td>
<td>•P</td>
<td>•S</td>
<td>•Cl</td>
<td>•Ar</td>
</tr>
<tr>
<td>•Na</td>
<td>•Ca</td>
<td>•Ga</td>
<td>•Ge</td>
<td>•As</td>
<td>•Se</td>
<td>•Br</td>
<td>•Kr</td>
</tr>
<tr>
<td>•K</td>
<td>•Sr</td>
<td>•In</td>
<td>•Sn</td>
<td>•Sb</td>
<td>•Te</td>
<td>•I</td>
<td>•Xe</td>
</tr>
<tr>
<td>•Rb</td>
<td>•Ba</td>
<td>•Tl</td>
<td>•Pb</td>
<td>•Bi</td>
<td>•Po</td>
<td>•At</td>
<td>•Rn</td>
</tr>
<tr>
<td>•Cs</td>
<td>•Ra</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

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Write Lewis dot symbols for the following:

(a) N

(b) S²⁻

(c) K⁺
Write Lewis dot symbols for the following:

(a) $N$

(b) $S^{2-}$

(c) $K^+$
8.2 Ionic Bonding

- Ionic bond: electrostatic force that holds oppositely charge particles together

- Formed between cations and anions

- Example

  \[
  \text{Na}^- + \text{Cl}^- \rightarrow \text{Na}^+ + \text{Cl}^-
  \]

  \[\text{iE}_1 + \text{EA}_1 = 496 \text{ kJ/mol} - 349 \text{ kJ/mol} = 147 \text{ kJ/mol}\]

  m.p. = 801°C  \[\Delta H_f^\circ = -410.9 \text{ kJ/mol}\]
Microscopic View of NaCl Formation

\[ \text{Cl}_2 + 2\text{Na} \rightarrow 2\text{NaCl} \]
• **Lattice energy** = the energy required to completely separate one mole of a solid ionic compound into gaseous ions

\[
\text{NaCl}(s) \rightarrow \text{Na}^+(g) + \text{Cl}^-(g) \quad \Delta H_{\text{lattice}} = +788 \text{ kJ/mol}
\]

Because they are defined as an *amount* of energy, lattice energies are always positive.
• Coulombic attraction:

\[ F \propto \frac{Q_1 \cdot Q_2}{d^2} \]

Q = amount of charge
d = distance of separation

• Lattice energy (like a coulombic force) depends on
  • Magnitude of charges
  • Distance between the charges
Lattice energies of alkali metal iodides

\[
F \propto \frac{(+1) \times (-1)}{(0.76 + 2.20)^2} \propto -0.11
\]

Largest lattice energy
(732 kJ/mol)

\[
F \propto \frac{(+1) \times (-1)}{(1.02 + 2.20)^2} \propto -0.10
\]

Intermediate lattice energy
(686 kJ/mol)

\[
F \propto \frac{(+1) \times (-1)}{(1.38 + 2.20)^2} \propto -0.08
\]

Smallest lattice energy
(632 kJ/mol)
The ionic radii sums for LiF and MgO are 2.01 and 2.06 Å, respectively, yet their lattice energies are 1030 and 3795 kJ/mol. Why is the lattice energy of MgO nearly four times that of LiF?
• **Born-Haber cycle**: A method to determine lattice energies
• Born-Haber cycle for CaO

\[
\begin{align*}
\text{Ca}(s) & \quad + \quad (1/2)\text{O}_2(g) & \quad \#6 \quad \rightarrow \quad \text{CaO}(s) \\
#1 & \quad \downarrow \quad \text{Ca}(g) & \quad \#3 & \quad \downarrow \quad \text{O}(g) \\
#2 & \quad \downarrow \quad \text{Ca}^{2+}(g) & \quad \#4 & \quad \downarrow \quad \text{O}^{2-}(g) & \quad \#5 \\
\end{align*}
\]

#1 Heat of sublimation = \( \Delta H_f^\circ[\text{Ca}(g)] = +178 \text{ kJ/mol} \)

#2 1st & 2nd ionization energies = \( I_1(\text{Ca}) + I_2(\text{Ca}) = +1734.5 \text{ kJ/mol} \)

#3 (1/2) Bond enthalpy = (1/2)D(O=O) = \( \Delta H_f^\circ[\text{O}(g)] = +247.5 \text{ kJ/mol} \)

#4 1st & 2nd electron affinities = \( E\text{A}_1(\text{O}) + E\text{A}_2(\text{O}) = +603 \text{ kJ/mol} \)

#5 \(-\) (Lattice Energy) = \(-\Delta H_{\text{lattice}}[\text{CaO}(s)] = \) (the unknown)

#6 Standard enthalpy of formation = \( \Delta H_f^\circ[\text{CaO}(s)] = -635 \text{ kJ/mol} \)

\[+178 +1734.5 +247.5 +603 -\Delta H_{\text{latt}} = -635\]

\[\Delta H_{\text{lattice}} = +3398 \text{ kJ/mol}\]
8.3 Covalent Bonding

- Atoms share electrons to form covalent bonds.

\[
\text{H} \cdot + \cdot \text{H} \rightarrow \text{H} \cdot \text{H} \quad \text{or} \quad \text{H} - \text{H}
\]

- In forming the bond the atoms achieve a more stable electron configuration.
• Octet: Eight is a “magic” number of electrons.

• Octet Rule: Atoms will gain, lose, or share electrons to acquire eight valence electrons

Examples:

\[
\text{Na}^- + \text{Cl}^- \rightarrow \text{Na}^+ + \text{Cl}^-
\]

\[
\text{H}^- + \text{H}^- + \text{O} \rightarrow \text{H}^-\text{O}^-\text{H}^+
\]


- Lewis Structures

\[
\begin{align*}
\text{H}^- + \text{H}^+ & \rightarrow \text{H}^+\text{H}^- \\
\text{Cl}^- + \text{Cl}^- & \rightarrow \text{Cl}^-\text{Cl}^- \\
\end{align*}
\]

Shared electrons $\Rightarrow$ Bonds
Non-bonding valence electrons $\Rightarrow$ Lone pairs
• Multiple Bonds

- The number of shared electron pairs is the number of bonds.

\[ \text{Cl–Cl} \quad \text{Cl–Cl} \quad \text{Single Bond} \]

\[ \text{O=C=O} \quad \text{O=C=O} \quad \text{Double Bond} \]

\[ \text{N≡N} \quad \text{N≡N} \quad \text{Triple Bond} \]
- **Bond strength and bond length**

  - **Bond strength**: single < double < triple
  - **Bond length**: single > double > triple

<table>
<thead>
<tr>
<th>Bond</th>
<th>Single</th>
<th>Double</th>
<th>Triple</th>
</tr>
</thead>
<tbody>
<tr>
<td>N–N</td>
<td>163 kJ/mol</td>
<td>418 kJ/mol</td>
<td>941 kJ/mol</td>
</tr>
<tr>
<td>N=N</td>
<td>1.47 Å</td>
<td>1.24 Å</td>
<td>1.10 Å</td>
</tr>
<tr>
<td>N≡N</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
8.4 Electronegativity and Polarity

- **Nonpolar covalent bond** = electrons are shared *equally* by two bonded atoms
- **Polar covalent bond** = electrons are shared *unequally* by two bonded atoms

<table>
<thead>
<tr>
<th>M:X</th>
<th>$M^\delta^+X^{\delta^-}$</th>
<th>$M^{+}X^{-}$</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Pure covalent bond</strong></td>
<td>Neutral atoms held together by <em>equally</em> shared electrons</td>
<td><strong>Polar covalent bond</strong></td>
</tr>
</tbody>
</table>
• Electron density distributions

- red ⇒ high electron density
- green ⇒ intermediate electron density
- blue ⇒ low electron density

 alternate representations

\[ \begin{align*}
\delta^+ & \quad \delta^- \\
H - F &
\end{align*} \]
• Electronegativity: ability of an atom to draw shared electrons to itself.

- More electronegative elements attract electrons more strongly.

  • relative scale
  • related to $IE$ and $EA$
  • unitless
  • smallest electronegativity: Cs 0.7
  • largest electronegativity: F 4.0
# Electronegativity: The Pauling Scale

Increasing electronegativity

<table>
<thead>
<tr>
<th>Period</th>
<th>Electronegativity</th>
<th>Elements</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>2.1</td>
<td>H, Li, Na, K, Rb, Cs, Fr</td>
</tr>
<tr>
<td>2</td>
<td>1.5</td>
<td>Be, Mg, Sc, Ti, V, Cr, Mn, Fe, Co, Ni, Cu, Zn, Ga, Ge, As, Se, Br, Kr</td>
</tr>
<tr>
<td>3</td>
<td>1.2</td>
<td>Al, Si, P, S, Cl, Ar</td>
</tr>
<tr>
<td>4</td>
<td>1.0</td>
<td>Ca, Sr, Y, Zr, Nb, Mo, Tc, Ru, Rh, Pd, Ag, Cd, In, Sn, Sb, Te, I, Xe</td>
</tr>
<tr>
<td>5</td>
<td>1.5</td>
<td>Sc, Ti, V, Cr, Mn, Fe, Co, Ni, Cu, Zn, Ga, Ge, As, Se, Br, Kr</td>
</tr>
<tr>
<td>6</td>
<td>1.2</td>
<td>Be, Mg, Sc, Ti, V, Cr, Mn, Fe, Co, Ni, Cu, Zn, Ga, Ge, As, Se, Br, Kr</td>
</tr>
<tr>
<td>7</td>
<td>1.0</td>
<td>Ca, Sr, Y, Zr, Nb, Mo, Tc, Ru, Rh, Pd, Ag, Cd, In, Sn, Sb, Te, I, Xe</td>
</tr>
</tbody>
</table>

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Variation in Electronegativity with Atomic Number
• Polar and nonpolar bonds

- \( H_2 \)
  - 2.1 - 2.1 = 0.0
  - nonpolar covalent

- \( HF \)
  - 4.0 - 2.1 = 1.9
  - polar covalent

- \( NaF \)
  - 4.0 - 0.9 = 3.1
  - ionic
  - > 2.0 is ionic
• Dipole moments and partial charges

- Polar bonds *often* result in polar molecules.
- A polar molecule possesses a *dipole*.

- **dipole moment** \((\mu)\) = the quantitative measure of a dipole

\[
\begin{align*}
\delta^+ & \quad \delta^- \\
H - F
\end{align*}
\]

\[
\mu = Qr
\]

SI unit: coulomb\(\cdot\)meter \((\text{C} \cdot \text{m})\)

common unit: debye \((\text{D})\)

\[1 \text{ D} = 3.34 \times 10^{-30} \text{ C} \cdot \text{m}\]

<table>
<thead>
<tr>
<th></th>
<th>(1.82 \text{ D})</th>
<th>(1.08 \text{ D})</th>
<th>(0.82 \text{ D})</th>
<th>(0.44 \text{ D})</th>
</tr>
</thead>
<tbody>
<tr>
<td>HF</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>HCl</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>HBr</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>HI</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
8.5 Drawing Lewis Structures

1) Draw skeletal structure with the central atom being the least electronegative element.

2) Sum the valence electrons. Add 1 electron for each negative charge and subtract 1 electron for each positive charge.

3) Subtract 2 electrons for each bond in the skeletal structure.

4) Complete electron octets for atoms bonded to the central atom except for hydrogen.

5) Place extra electrons on the central atom.

6) Add multiple bonds if atoms lack an octet.
What is the Lewis structure of NO$_3^-$?

1) Draw skeletal structure with central atom being the least electronegative.

2) Sum valence electrons. Add 1 for each negative charge and subtract 1 for each positive charge.

$$\text{NO}_3^- \Rightarrow (1 \times 5) + (3 \times 6) + 1 = 24 \text{ valence } e^-$$

3) Subtract 2 for each bond in the skeletal structure.

4) Complete electron octets for atoms bonded to the central atom except for hydrogen.

5) Place extra electrons on the central atom.

6) Add multiple bonds if atoms lack an octet.
<table>
<thead>
<tr>
<th>Step</th>
<th>CH₄</th>
<th>CCl₄</th>
<th>H₂O</th>
<th>O₂</th>
<th>CN⁻</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>H–C–H</td>
<td>Cl–C–Cl</td>
<td>H–O–H</td>
<td>O–O</td>
<td>C–N</td>
</tr>
<tr>
<td>2</td>
<td>8</td>
<td>32</td>
<td>8</td>
<td>12</td>
<td>10</td>
</tr>
<tr>
<td>3</td>
<td>8 – 8 = 0</td>
<td>32 – 8 = 24</td>
<td>8 – 4 = 4</td>
<td>12 – 2 = 10</td>
<td>10 – 2 = 8</td>
</tr>
<tr>
<td>5</td>
<td>—</td>
<td>—</td>
<td>H–O–H</td>
<td>—</td>
<td>—</td>
</tr>
<tr>
<td>6</td>
<td>—</td>
<td>—</td>
<td>—</td>
<td>:O≡O⁻</td>
<td>[C≡N⁻]</td>
</tr>
</tbody>
</table>
8.6 Lewis Structures and Formal Charge

• The electron surplus or deficit, relative to the free atom, that is assigned to an atom in a Lewis structure.

Formal Charge = Total valence electrons - Total non-bonding electrons - \( \frac{1}{2} \) bonding electrons

Example: \( \text{H}_2\text{O} = \text{H} : \overset{\cdot}{\text{O}} : \text{H} \)

H: orig. valence \( e^- = 1 \)
- non-bonding \( e^- = 0 \)
- 1/2 bonding \( e^- = -1 \)
formal charge = 0

O: orig. valence \( e^- = 6 \)
- non-bonding \( e^- = -4 \)
- 1/2 bonding \( e^- = -2 \)
formal charge = 0

Formal charges are not “real” charges.
Example: Formal charges on the atoms in ozone

\[
\begin{align*}
O &= 6 - 4 - \frac{1}{2}(4) \\
    &= 0 \\
O &= 6 - 2 - \frac{1}{2}(6) \\
    &= +1 \\
O &= 6 - 6 - \frac{1}{2}(2) \\
    &= -1
\end{align*}
\]
• Formal charge guidelines
  - A Lewis structure with *no* formal charges is generally better than one with formal charges.
  - Small formal charges are generally better than large formal charges.
  - Negative formal charges should be on the more electronegative atom(s).

Example:

```
H   C   O   H or C   O   ?
```

Answer:

```
H−\overset{−}{\underset{\text{H}}{C==O}}−H
```

```
H
\overset{\text{C==O}}{C==O}\overset{\text{H}}{H}
```
Identify the best structure for the isocyanate ion below:

(a) \([\ddots C = N = O \ddots ]^–\)  
   \(-2 \quad +1 \quad 0\)

(b) \([\ddots C \equiv N - O \ddots ]^–\)  
   \(-1 \quad +1 \quad -1\)

(c) \([\ddots C - N \equiv O \ddots ]^–\)  
   \(-3 \quad +1 \quad +1\)
Identify the best structure for the isocyanate ion below:

(a) \[ :\overset{\cdots}{C} = \overset{\cdots}{N} = \overset{\cdots}{O}: ]^− \]
\[ \begin{array}{ccc}
-2 & +1 & 0 \\
\end{array} \]

(b) \[ :\overset{\cdots}{C} \equiv \overset{\cdots}{N} - \overset{\cdots}{O}: ]^− \]
\[ \begin{array}{ccc}
-1 & +1 & -1 \\
\end{array} \]

(c) \[ :\overset{\cdots}{C} - \overset{\cdots}{N} \equiv \overset{\cdots}{O}: ]^− \]
\[ \begin{array}{ccc}
-3 & +1 & +1 \\
\end{array} \]
8.7 Resonance

- **Resonance** structures are used when two or more equally valid Lewis structures can be written.

Example: $\text{NO}_2$

$\left[\begin{array}{c} \text{O} \rightarrow \text{N} = \text{O} : \\
\end{array}\right]$

These two bonds are known to be identical.

Solution: $\left[\begin{array}{c} \text{:O} \rightarrow \text{N} = \text{O} : \\
\end{array}\right] \leftrightarrow \left[\begin{array}{c} \text{O} \rightarrow \text{N} \rightarrow \text{O} : \\
\end{array}\right]$

Two **resonance structures**, their average or the **resonance hybrid**, best describes the nitrite ion.

The double-headed arrow indicates resonance.
Additional Examples

Carbonate: $\text{CO}_3^{2-}$

$\begin{bmatrix} \cdot \overset{\bullet}{-} \cdot \\ \cdot \overset{-}{-} \cdot \overset{-}{-} \cdot \overset{-}{-} \cdot \overset{-}{-} \cdot \overset{-}{-} \cdot \overset{-}{-} \cdot \overset{-}{-} \end{bmatrix}^{2-}$

Benzene: $\text{C}_6\text{H}_6$

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Exceptions to the octet rule fall into three categories:

- Molecules with an incomplete octet
- Molecules with an odd number of electrons
- Molecules with an expanded octet
• Incomplete Octets

Example: BF$_3$ (boron trifluoride)

\[ BF_3 \Rightarrow (1 \times 3) + (3 \times 7) = 24 \text{ val. } e^- \]

\[ \vdots F: \]
\[ \vdots F - B - F: \]
\[ \vdots F: \]
\[ \vdots F - B = F:+1 \]

- Common with Be, B and Al compounds, but they often dimerize or polymerize.

\[ \text{Example: } \begin{array}{c}
\text{Cl} \quad \text{Be} \\
\text{Cl} \\
\text{Cl} \\
\text{Be} \\
\text{Cl} \quad \text{Be} \\
\text{Cl} \\
\text{Cl} \\
\end{array} \]
• Odd Numbers of Electrons

Example: NO (nitrogen monoxide or nitric oxide)

\[
\text{NO} \Rightarrow (1 \times 5) + (1 \times 6) = 11 \text{ valence } e^- \\
\left\{ \begin{array}{l}
0 \\
0 \\
N = O:
\end{array} \right. \leftrightarrow \left\{ \begin{array}{l}
0 \\
0 \\
N = O:
\end{array} \right.
\]

Are these both equally good?

Example: NO\textsubscript{2} (nitrogen dioxide)

\[
\text{NO}\textsubscript{2} \Rightarrow (1 \times 5) + (2 \times 6) = 17 \text{ val. } e^- \\
\left\{ \begin{array}{l}
0 \\
0 \\
0 \\
O = N - O:
\end{array} \right. \leftrightarrow \left\{ \begin{array}{l}
0 \\
0 \\
0 \\
O - N = O:
\end{array} \right. \leftrightarrow \left\{ \begin{array}{l}
0 \\
0 \\
0 \\
O = N - O:
\end{array} \right. \leftrightarrow \left\{ \begin{array}{l}
0 \\
0 \\
0 \\
O - N = O:
\end{array} \right.
\]

Are these all equally good?
• Expanded Octet

- Elements of the 3rd period and beyond have $d$-orbitals that allow more than 8 valence electrons.

$\text{SF}_6 = \begin{array}{c}
\text{S} \\
\text{F} & \text{F} & \text{F} & \text{F} & \text{F} & \text{F}
\end{array}$

48 valence e$^-$(S has 12 valence electrons)

$\text{XeF}_2 = \begin{array}{c}
\text{F} & \text{Xe} & \text{F}
\end{array}$

22 valence e$^-$(Xe has 10 valence electrons)
8.9 Bond Enthalpy

- Bond enthalpy is the energy associated with breaking a particular bond in one mole of gaseous molecules.
  - Bond enthalpy is one measure of molecular stability.
  - Symbol: $\Delta H^\circ$
  - For diatomic molecules these are accurately measured quantities.

$$\text{Cl}_2(g) \rightarrow \text{Cl}(g) + \text{Cl}(g) \quad \Delta H^\circ = 243.4 \text{ kJ}$$

$$\text{HCl}(g) \rightarrow \text{H}(g) + \text{Cl}(g) \quad \Delta H^\circ = 431.9 \text{ kJ}$$

$$\text{O}_2(g) \rightarrow \text{O}(g) + \text{O}(g) \quad \Delta H^\circ = 495.0 \text{ kJ}$$

- single bonds

$$\text{N}_2(g) \rightarrow \text{N}(g) + \text{N}(g) \quad \Delta H^\circ = 945.4 \text{ kJ}$$

- double bond

- triple bond
- Bond enthalpies for polyatomic molecules depend upon the bond’s environment.

\[ \text{H–C–H} \rightarrow \text{H–C} + \text{H} \quad \Delta H = 435 \text{ kJ} \]

\[ \text{H–C–C–H} \rightarrow \text{H–C–C} + \text{H} \quad \Delta H = 410 \text{ kJ} \quad 6\% \text{ less} \]

- Average bond enthalpies are used for polyatomic molecules.

• Provide only estimates
• Prediction of bond enthalpy

\[ \Delta H^o = \Sigma \text{BE(reactants)} - \Sigma \text{BE(products)} \]
Example: Calculate the enthalpy of reaction for

\[
CH_4(g) + Br_2(g) \rightarrow CH_3Br(g) + HBr(g)
\]

Solution: Consider ONLY bonds broken or formed.

\[\Delta H_{rxn} = [BE(C-H) + BE(\text{Br-Br})] - [BE(C-\text{Br}) + BE(H-\text{Br})] \]

\[= [(413) + (193)] - [(276) + (366)] \]

\[= -36 \text{ kJ/mol} \]
<table>
<thead>
<tr>
<th>Bond</th>
<th>Bond Enthalpy (kJ/mol)</th>
<th>Bond</th>
<th>Bond Enthalpy (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>H–H*</td>
<td>436.4</td>
<td>C–S</td>
<td>255</td>
</tr>
<tr>
<td>H–N</td>
<td>393</td>
<td>C=O</td>
<td>477</td>
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<tr>
<td>H–O</td>
<td>460</td>
<td>N–N</td>
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</tr>
<tr>
<td>H–S</td>
<td>368</td>
<td>N≡N</td>
<td>418</td>
</tr>
<tr>
<td>H–P</td>
<td>326</td>
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<td>941.4</td>
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<td>H–F</td>
<td>568.2</td>
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<td>176</td>
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<td>H–Cl</td>
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<td>C≡C</td>
<td>812</td>
<td>P≡P</td>
<td>489</td>
</tr>
<tr>
<td>C–N</td>
<td>276</td>
<td>S–S</td>
<td>268</td>
</tr>
<tr>
<td>C=N</td>
<td>615</td>
<td>S=O</td>
<td>352</td>
</tr>
<tr>
<td>C≡N</td>
<td>891</td>
<td>F–F</td>
<td>156.9</td>
</tr>
<tr>
<td>C–O</td>
<td>351</td>
<td>Cl–Cl</td>
<td>242.7</td>
</tr>
<tr>
<td>C=O↑</td>
<td>745</td>
<td>Cl–F</td>
<td>193</td>
</tr>
<tr>
<td>C≡O</td>
<td>1070</td>
<td>Br–Br</td>
<td>192.5</td>
</tr>
<tr>
<td>C–P</td>
<td>263</td>
<td>I–I</td>
<td>151.0</td>
</tr>
</tbody>
</table>

*Bond enthalpies shown in red are for diatomic molecules.

†The C≡O bond enthalpy in CO₂ is 799 kJ/mol.
Key Points

• Lewis dot symbols
• Ionic bonding
• Lattice energy
• Born-Haber cycle
• Covalent bonding
• Octet rule
• Lewis structures
• Bond order
• Bond polarity
Key Points

- Electronegativity
- Dipole moment
- Drawing lewis structures
- Formal charge
- Resonance structures
- Incomplete octets
- Odd numbers of electrons
- Expanded octets
- Bond enthalpy