

Essential Question

How does energy change as two atoms form a bond?

Forces Between Two Atoms

- Each nucleus attracts the other atom's electrons.
- The two nuclei repel each other.
- The electrons repel each other.

Opposite charges attract
and like charges repel.

Potential Energy vs. Distance

- The graph shows how potential energy (PE) changes as two atoms approach each other.
- The x-axis is the **internuclear distance**, the distance between the centers of the two nuclei (pm). The y-axis is potential energy (kJ/mol).

Potential Energy vs. Distance

- Far apart: the atoms don't interact, so $PE = 0$.
- Lower PE = more stable.

Reading the Curve

1. **Far apart:** no attraction and no repulsion, so PE is about zero.
2. **Closer:** attraction is stronger than repulsion, so PE drops.
3. **Too close:** the nuclei repel strongly, so repulsion wins and PE rises sharply.

Bond Length and Bond Energy

- **Bond length:** the internuclear distance at the lowest PE.
- **Bond energy:** the energy needed to separate the atoms, measured from zero down to the bottom of the well.

Exam tip: a deeper well means a stronger bond.

Bond Order Affects Bond Length

- Higher bond order → shorter, stronger bond.
- More shared electrons between the nuclei means a stronger attraction that pulls the nuclei closer together.
- A double bond is not exactly twice as strong as a single bond (614 kJ/mol, not $2 \times 347 = 694$ kJ/mol), so don't just multiply.

Bond order is the number of shared electron pairs between two atoms: 1 for a single bond, 2 for a double bond, and 3 for a triple bond.

Bond Order on a PE Graph

- Farther left → shorter bond
- Deeper well → stronger bond

The $\text{C}\equiv\text{C}$ triple bond's minimum is farther left (shorter bond) and its well is much deeper (stronger bond) than the $\text{C}-\text{C}$ single bond's.

Atom Size Affects Bond Length

- Bigger atoms → longer bonds
- Longer bonds → usually weaker

Going down the halogen group, the atoms get bigger because they have more electron shells (a larger atomic core).

Coulomb's Law

- q_1 = charge of the cation
- q_2 = charge of the anion
- r = distance between the nuclei (center to center)

$$F \propto q_1 q_2 / r^2$$

Larger Charges, Stronger Attraction

- Similar distance (r)
- Larger charges $\rightarrow$ stronger attraction

NaF and CaO have almost the same r (about 235 pm and 240 pm).

Smaller Ions, Stronger Attraction

- Same charges
- Smaller ions $\rightarrow$ smaller r $\rightarrow$ stronger attraction

NaF and NaI both have charges of +1 and -1.

Example: Reading a PE Graph

Try it first!

The graph shows potential energy vs. internuclear distance for Cl_2 . Estimate the bond length and the bond energy.

Diagram: Graph of potential energy (kJ/mol, -300 to 300) versus internuclear distance (pm, 0 to 500) for Cl_2 , with gridlines every 50 units. The curve drops from above 300 kJ/mol to a minimum near 200 pm, then rises toward zero.

Example: Comparing Two Bonds

Try it first!

Curves X and Y show N_2 and F_2 . Which curve represents N_2 ? Justify your answer.

Diagram: Two potential energy curves. Curve X (green) has a very deep well with its minimum farther left. Curve Y (red) has a shallow well with its minimum farther right.

Example: Using Coulomb's Law

Try it first!

Use Coulomb's law to estimate how many times stronger the attraction is in MgO than in NaCl. Ionic radii: $\text{Na}^+ = 102 \text{ pm}$, $\text{Cl}^- = 181 \text{ pm}$, $\text{Mg}^{2+} = 72 \text{ pm}$, $\text{O}^{2-} = 140 \text{ pm}$.

Big Ideas

- Lowest PE $\rightarrow$ bond length.
- Well depth $\rightarrow$ bond energy.
- Higher bond order $\rightarrow$ shorter, stronger bond.
- Bigger charges and smaller ions $\rightarrow$ stronger attraction.