Whenever a question asks you to look at bond energy, bond leghth, bond strength or bond order – then look inside the molecule (not at IMF)

Bond energy: the energy needed to break one mole a compound's molecules into separate atoms

Bond length: The average distance between the nuclei of two bonded atoms

<u>Relationship between bond strength and</u>

- bond length
- atom/molecule size
- bond order (no of bond between atoms)

It is important to remember that bond length is measured between two atoms that are bonded to each other. The following diagrams show the bond length for CO and for CO_2. The grey circle represents carbon and the white circle represents oxygen.

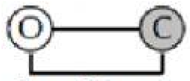


Figure 3.10:
The bond length for carbon monoxide (CO).

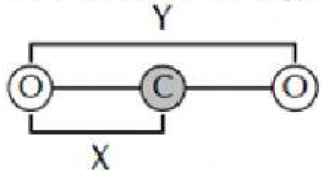


Figure 3.11:
The bond length for each C – O bond in carbon dioxide (CO_2) is indicated by X. **Y is not** the bond length.

A third property of bonds is the bond strength. **Bond strength** means how strongly one atom attracts and is held to another. The strength of a bond is related to the bond length, the size of the bonded atoms and the number of bonds between the atoms. In general:

- the shorter the bond length, the   the bond between the atoms.
- the smaller the atoms involved, the   the bond.
- the more bonds that exist between the same atoms, the   the bond.

  (double and triple bonds are stronger than single bonds)

- The more bond energy required the _______ the bond

Relationship between bond length and bond order

As a general trend, bond length decreases across a period in the periodic table and increases down a group (think about atomic radius from gr 10).

Atoms with multiple bonds between them have shorter bond lengths than singly bonded ones.

BOND	LENGTH (pm)	ENERGY kJ/mol
O=O	121	498
O–O	148	145

Examples

1. Write the following 3 substances in order to increasing boiling point: HF, HBr and HCl

If the question asks what forces are needed to be overcome to make a substance boil- then it is asking what intermolecular forces need to be overcome to change the phase. Then you focus on the type of intermolecular forces. If the IMF between the substances that are being compared are the same- then look at the atomic radius of the substance.

Atomic radius increases as you move down a group and decreases from left to right across a period.

__________ __________

HF has the strongest IMF (hydrogen bonding) between them.

Then HBr and HCl have the same Van der Waals, dipole-dipole forces between them.

Now look at molecular size.

In terms of size HBr is bigger than HCl (remember atomic radius increases as you move down a group) So the bigger the molecule the higher the boiling point.

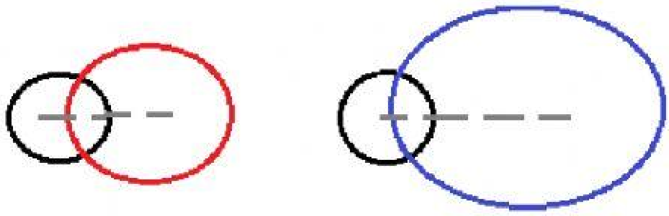
Thus HF has the highest boiling point, then HCl then HBr.

Bond energy/length/strength all focus on the inside of the actual molecule – not the IMF!

2. Compare the **bond length** of HF and HBr

Bond length is defined as the distance between the centres of two covalently bonded atoms. Generally, the length of the bond between two atoms is approximately the sum of the covalent radii of the two atoms.

HF will have a shorter bond length than HBr, since F has a smaller atomic radius



3. Compare the **bond energy** of HF and HBr

Bond energy

The shorter the bond length – the higher the bond energy (atoms are closer together and more difficult to break apart.) Another factor here is also the EN difference. HF has a bigger EN difference and is thus a more polar molecule. Thus is has a stronger bond and higher bond energy.

Question comparing different molecules

Question on melting point, boiling point, vapour pressure or viscosity

Focus on differences in IMF

The stronger the IMF- the higher the BP etc

(if any ionic molecules are included – there are very strong ionic forces between them)

If **IMF strength** is the same then you look at:

IMF force strength increases with an increase in molecular mass.

Question on bond length

Focus on atomic radius

The greater the radius the greater the bond length (and weaker forces between atoms) and thus smaller bond energy

Question on bond energy

Focus on intramolecular force strength, then bond length, size of atoms and no of bonds

The stronger the intramolecular strength the higher the bond energy

If IMF strength or **intramolecular bonds** are the same then you look at:

EN difference and polarity.

If a molecule has a bigger EN difference then it is more polar and has higher BP or bond energy

Solubility

Like dissolves like

Polar substances in dissolve in polar solvents

Non-polar substances dissolve in non-polar substances

Or substances with similar IMF strength's between molecules are miscible in each other.

Why do some atoms bond and form molecules while other do not?

Case 1: A bond forms

Let's start by imagining that there are two hydrogen atoms approaching one another. As they move closer together, there are three forces that act on the atoms at the same time. These forces are described below:

1. **repulsive force** between the electrons of the atoms, since like charges repel

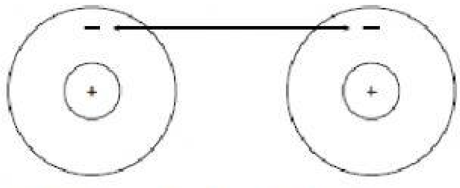


Figure 3.2: Repulsion between electrons

2. **attractive force** between the nucleus of one atom and the electrons of another

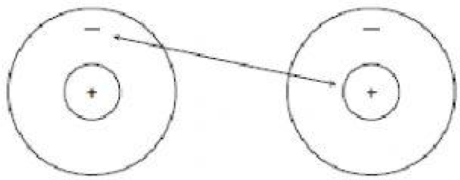


Figure 3.3: Attraction between electrons and protons.

3. **repulsive force** between the two positively-charged nuclei

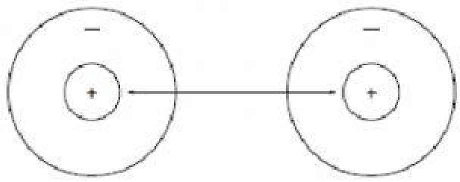
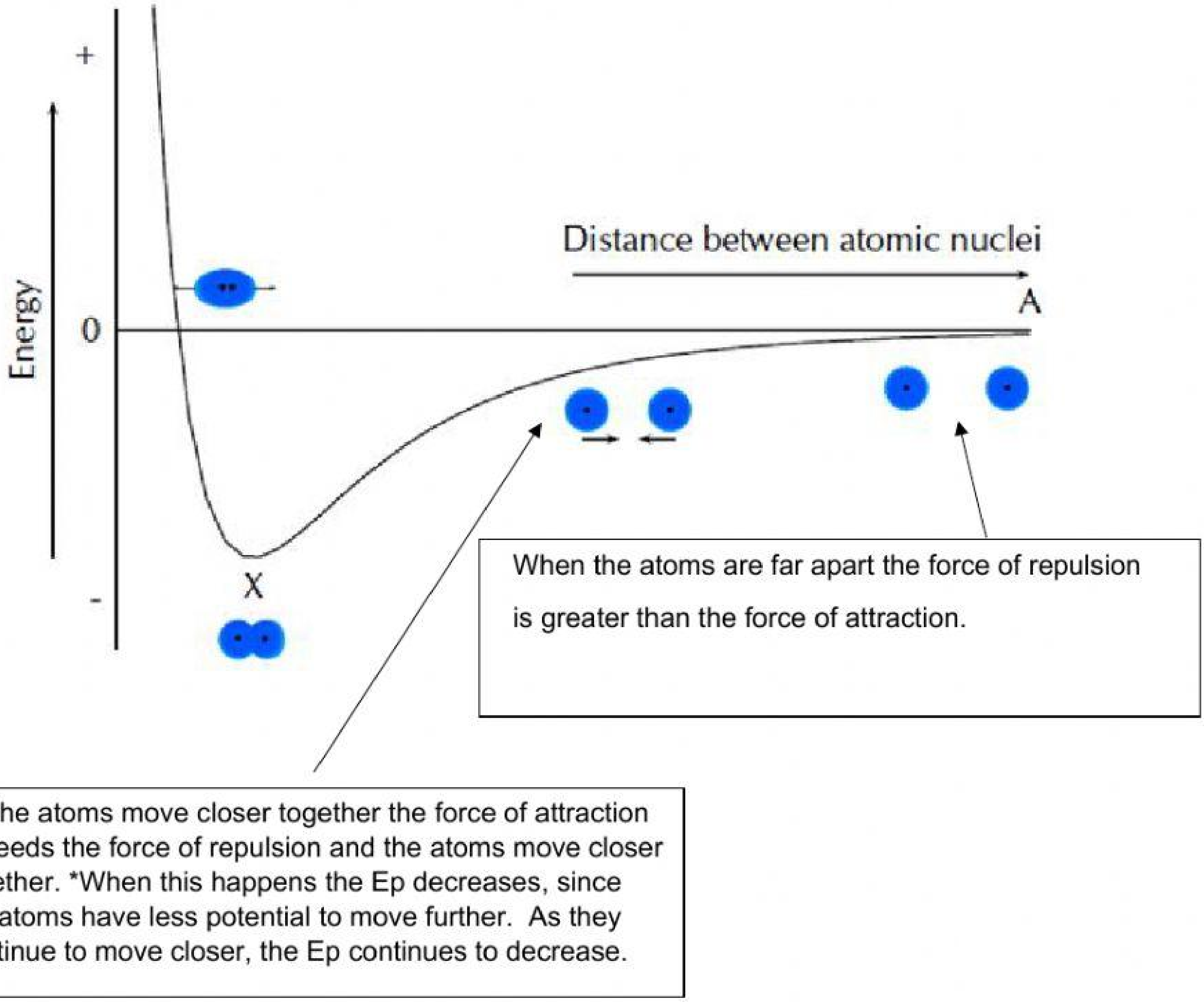


Figure 3.4: Repulsion between protons

These three forces work together when two atoms come close together. As the total force experienced by the atoms changes, the amount of energy in the system also changes.

Now look at Figure 3.5 to understand the energy changes that take place when the two atoms move towards each other.

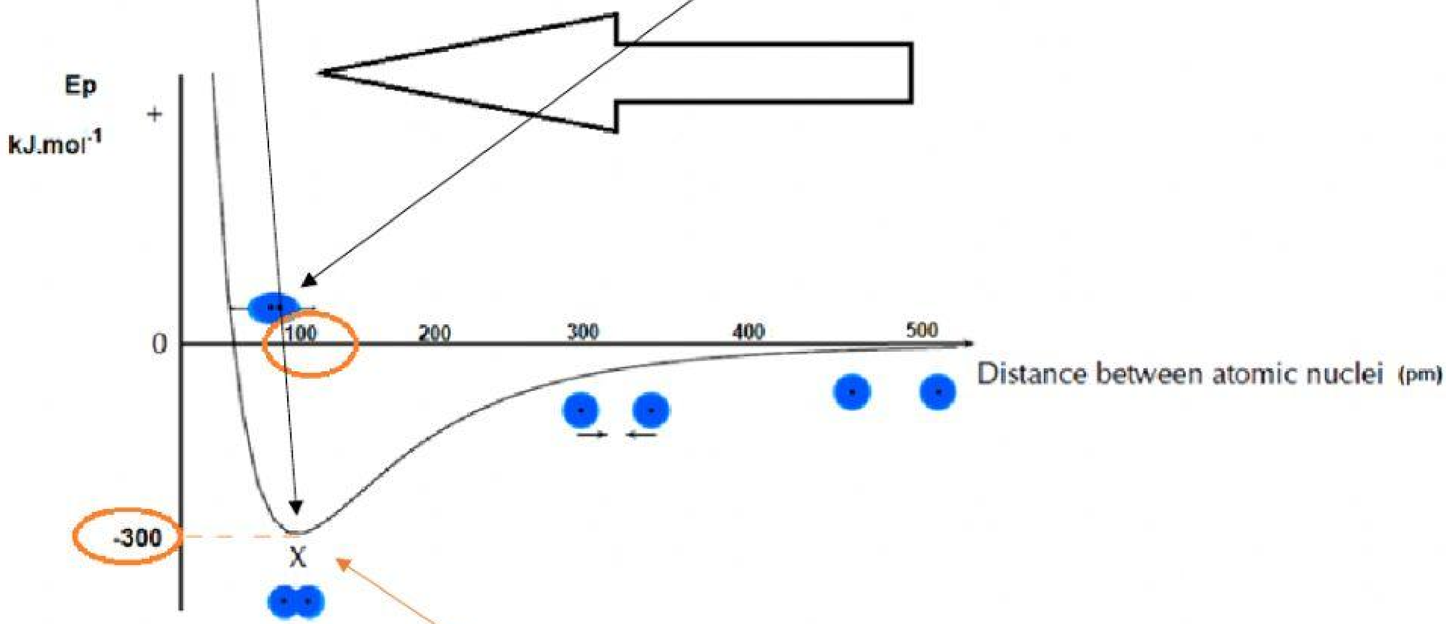
This graph is best understood when read from right to left.



The Ep is at it's lowest at point X, since the atoms have bonded together and form a stable compound. They no longer want to move.

If the atoms were to be forced closer together then the Ep would increase drastically. But why?

If the atoms are forced closer then the positive nuclei are forced closer together. These nuclei would desperately want to repel each other and thus the atoms would desperately want to move apart and thus have a very high Ep.



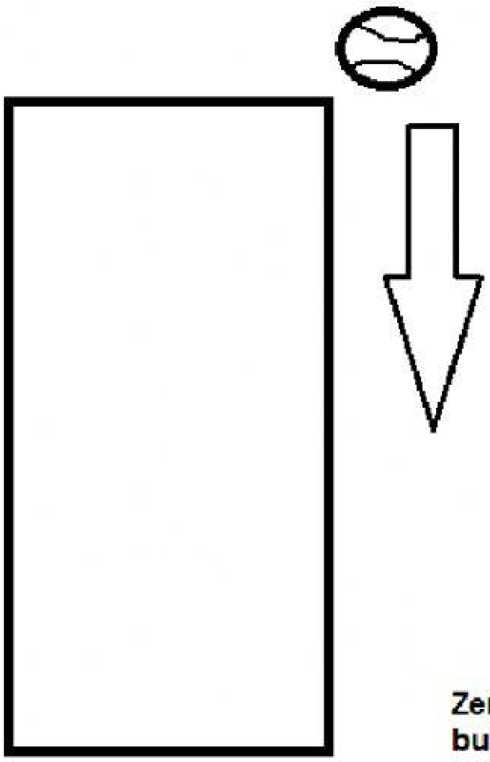
This is a very significant point, as two things can be read from the graph here

The bond energy is -300kJ.mol^{-1} (remember this is the energy needed to break a bond in a molecule)

And the bond length in this molecule is 100pm (picometers)

(This is the average distance between the nuclei of two bonded atoms)

Compare the Ep to that of a ball that is dropped from a high building



Max Ep since ball has the most potential energy (it has the furthest distance to move from here)

Zero Ep when it reaches the bottom of the building since the ball has no more potential to move. It has reached it's destination.

More in depth explanation of the change in Ep

Let us imagine that we have fixed the one atom and we will move the other atom closer to the first atom. As we move the second hydrogen atom closer to the first (from point A to point X) the energy of the system decreases. Attractive forces dominate this part of the interaction. As the second atom approaches the first one and gets closer to point X, more energy is needed to pull the atoms apart. This gives a negative potential energy.

At point X, the attractive and repulsive forces acting on the two hydrogen atoms are balanced. The energy of the system is at a minimum.

Further to the left of point X, the repulsive forces are stronger than the attractive forces and the energy of the system increases.

For hydrogen the energy at point X is low enough that the two atoms stay together and do not break apart again. This is why when we draw the Lewis diagram for a hydrogen molecule we draw two hydrogen atoms next to each other with an electron pair between them.

$$\mathbf{H} \overset{\times}{\bullet} \mathbf{H}$$

We also note that this arrangement gives both hydrogen atoms a full outermost energy level (through the sharing of electrons or covalent bonding).