The term dipole moment can be defined either with respect to chemical bonds or with respect to molecules.

**Bond Dipole Moment**

The electrons in a covalent bond connecting two different atoms are not equally shared by the atoms due to the electronegativity difference between the two elements. The atom of the more electronegative element has the greater share of the electrons than the atom of the less electronegative element. Consequently, the atom that has the greater share of the bonding electrons bears a partial negative charge and the atom that has the lesser share a partial positive charge of equal magnitude. For example, consider the hypothetical diatomic molecule AB. Assume that element B is more electronegative than element A. Atom B has the greater share of the electrons in the AB bond. That is, AB bond is polarized toward atom B, and there is a partial negative charge (-δ) on atom B and a partial positive charge of equal magnitude (+δ) on atom A.

\[
\text{A} \quad \rightarrow \quad \text{B}
\]

The product of the magnitude of the charge either on atom A or on atom B (δ) and AB bond length (d) is called the dipole moment (symbol: \(\mu\)) of the bond.

\[
\mu \ (\text{AB}) = \delta \times d
\]

The SI units of dipole moment are debye (D).

**2. Molecular Dipole Moment**

By definition, molecules are electrically neutral. However, many a molecule has a region, or pole, bearing a net partial positive charge, and the opposite region, or pole, bears a net partial negative charge of equal magnitude, as shown below schematically.

Such a molecule has a dipole moment, which is equal to the vector sum of the dipole moments of all bonds in the molecule. The dipole moment of a molecule, being a vector, has a direction, which is shown using the following symbol.
A molecule may not have a dipole moment despite containing bonds that do. For example, each of the two carbon-oxygen bonds in CO$_2$ has a dipole moment, but the CO$_2$ molecule has no dipole moment because the dipole moments of the two carbon-oxygen bonds are identical in magnitude and opposite in direction, resulting in a vector sum of zero.

Bond dipole moments can not be determined experimentally; only molecular dipole moments can. However, there are indirect ways to determine bond dipole moments, approximately if not accurately. A diatomic molecule has only one bond. Therefore, the dipole moment of the bond in a diatomic molecule is equal to the dipole moment of the molecule. For example, experimentally, the dipole moment of HF is 1.91 D. Thus, the dipole moment of the hydrogen-fluoride bond in HF is 1.91 D. In some organic molecules, one bond contributes overwhelmingly to the molecular dipole moment, making the contribution of the other bonds insignificant. Consequently, the dipole moment of the bond that contributes to the molecular dipole moment is approximately equal to the molecular dipole moment. For example, consider CH$_3$Cl.
• electronegativity difference: C and Cl >>> C and H
• bond length: C-Cl >>> C-H

Thus, dipole moment: C-Cl >>> C-H and the dipole moment of C-Cl is approximately the same as that of CH₃Cl. Experimentally, the dipole moment of CH₃Cl is 1.87 D, meaning the dipole moment of the C-Cl bond in CH₃Cl is approximately 1.87 D.

Molecular dipole moment can provide useful information regarding the structure of the molecule. For example, the dipole moment of water is 1.85 D, implying that the water molecule is not linear (1), for the dipole moment of a linear water molecule would be zero.

Contributors
• Gamini Gunawardena from the OChemPal site (Utah Valley University)