Chemical Bonds:

Ionic compounds formed from the transfer of an electron. Electrostatic attraction.

Na [Ne]$3s^1$
Cl [Ne]$3s^2 3p^5$

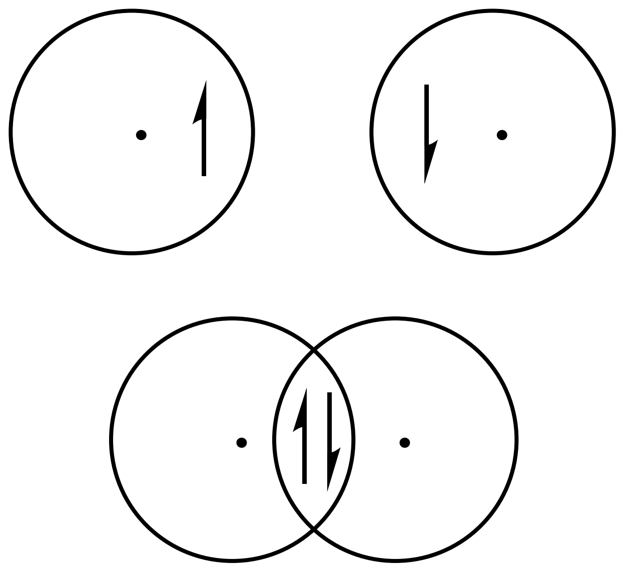
Na^+ [Ne]$3s^0$ or [Ne]
Cl^- [Ne]$3s^2 3p^6$ or [Ar]

Covalent Bond:

Covalent bond is a bond formed between two atoms when they share electrons.

Consider H_2 molecule.

H: $1s^1$ **$1s^1$:H**



H—H

Lewis Structures:

Method of representing molecular structures using the symbol of the element to represent the atom and dots(·) for the valence electrons and dashes(-) for shared electrons.

The “ - ” represents a shared pair(2 electrons) between atoms.

Single covalent bond (–) involves the sharing of 2 valence electrons between atoms.

Double covalent bond (=) involves the sharing of 4 valence electrons between atoms.

Electronegativity:

A relative measure of an atoms ability in a molecule to attract electrons to itself.

Non-polar Bond-Equal sharing of electrons between atoms.

Polar Bond- Unequal sharing of electrons between atoms.

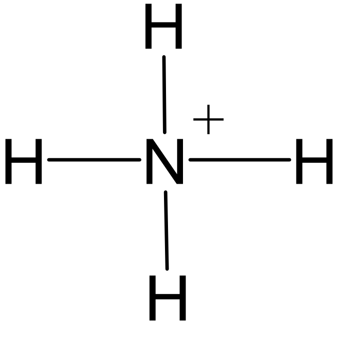
Ionic Bond-Transfer of electrons between atoms.

Formal Charges:

Formal charges are apparent charges associated with some atoms in lewis structures.

Formal Charge = (# valence electrons) - (# bonds) - (#unshared electrons)

Ex:



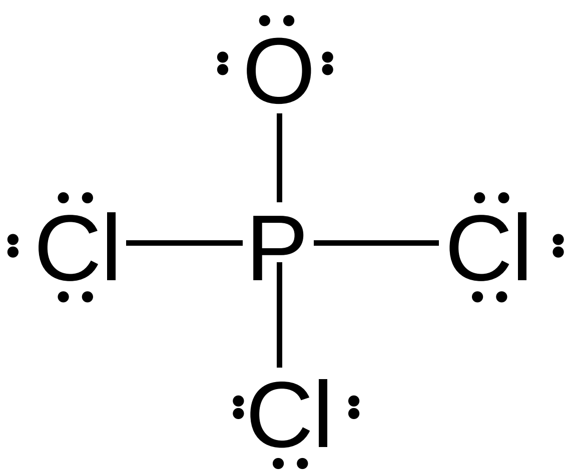
Ex: NH_4^+

N: formal charge = 5 - 4 - 0 = 5 - 4 = +1

H: formal charge = 1 - 1 - 0 = 1 - 1 = 0

Ex:

Determine the formal charge for the following lewis structure.



Drawing Lewis Structures:

Determine the lewis structure for ClO_3^-.

STEP 1: Find the total number of valence electrons supplied by all the atoms in the structure. If an anion, increase the total number of valence electrons by the charge. If a cation, decrease the total number of valence electrons by the charge.

STEP 2: Add up the total # of valence electrons required. There are 2 needed for each H atom and 8 for other elements.

STEP 3: The # in STEP 2 minus that in STEP 1 is the number of electrons that must be shared. Divide this number by 2 to obtain the total number of bonds.

STEP 4: Write out the elements with the one with the lowest electronegativity in the centre.

STEP 5: Indicate covalent bonds using dashes. Connect each atom and then add any multiple bonds until the # obtained is equal to the # of bonds calculated in STEP 3.

STEP 6: The total # of electrons from STEP 1 minus the # bonding electrons in STEP 3 is equal to the # of unshared electrons. Complete the octet of each atom.

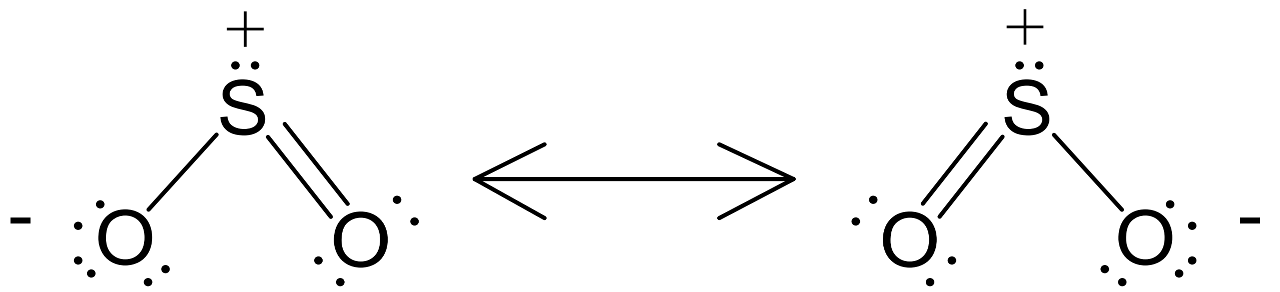
STEP 7: Indicate the formal charges.

Ex:2

Draw the lewis structure of SO_2.

Resonance:

Occurs when a molecule or ion has two or more possible lewis structures.



Lewis Theory is limited.

Ex:

Draw the resonance forms for CO_3^{2-}.

1. Resonance structures can not have adjacent atoms of like formal charges.
2. Most stable resonance structure has the fewest formal charges and formal charges of the lowest magnitude.

Ex:2
Draw the resonance forms for FNO_2.

Ex:3
Draw the resonance forms for OCN^-.

Bond Dissociation Energy:

The energy required to break the bond holding two atoms together in a diatomic molecule is called the bond dissociation energy.

$$H\text{–}H(g) \rightarrow 2H(g) \qquad \Delta H = +435\ kJ$$

$$Cl\text{–}Cl(g) \rightarrow 2Cl(g) \qquad \Delta H = +243\ kJ$$

It requires energy to break a covalent bond.

$$\text{H–Cl(g)} \rightarrow \text{H(g)} + \text{Cl(g)} \quad \Delta H = +431 \text{ kJ}$$

Thus,

$$\text{H(g)} + \text{Cl(g)} \rightarrow \text{H–Cl(g)} \quad \Delta H = -431 \text{ kJ}$$

When a bond is formed, energy is given off.

ΔH = Energy to break bonds - Energy Released on Bond Formation

When a chemical reaction occurs, the heat given off is due to the breaking and forming of chemical bonds. Heat change observed(ΔH) can be related to the Bond Dissociation Energy.

Ex:

$$\mathbf{H{-}H(g) + Cl{-}Cl(g) \rightarrow 2H{-}Cl(g) \quad \Delta H = -184.6\ kJ}$$

$$\mathbf{H{-}H(g) \rightarrow 2H(g) \quad \Delta H = +435\ kJ}$$

$$\mathbf{Cl{-}Cl(g) \rightarrow 2Cl(g) \quad \Delta H = +243\ kJ}$$

$$\mathbf{\underline{2H(g) + 2Cl(g) \rightarrow 2H{-}Cl(g) \quad \Delta H = 2\times{-431}\ kJ}}$$

$$\mathbf{H{-}H(g) + Cl{-}Cl(g) \rightarrow 2H{-}Cl(g) \quad \Delta H = -184.6\ kJ}$$