8.4 The Properties of Ionic Compounds

- Ionic solids – crystalline solids (regular three-dimensional arrays of stacked ions)
  - high melting and boiling points – very strong attractions between the ions (hard to separate)
  - hard and brittle
  - some are soluble (electrolytes)

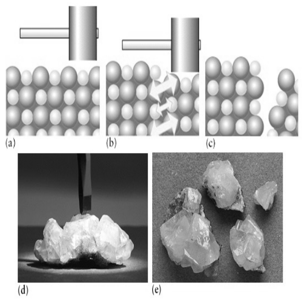


Fig. 8.11

Covalent Bonds

8.5 From Atoms to Molecules

- **Covalent bond** – a result of atoms sharing a pair of electrons

  H· + ·H → H:H

  - the negative charge of the shared pair of **e**$^-$ attracts the two positive nuclei
  - the electron density between the nuclei increases

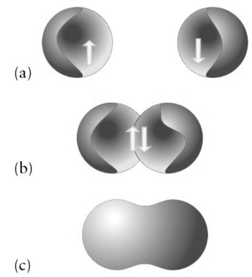
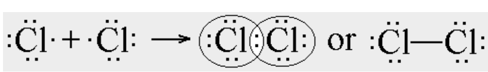


Fig. 8.13

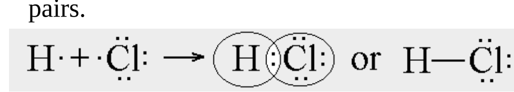
8.6 The Octet Rule and Lewis Structures

- **Octet rule** – in covalent bonding atoms share pairs of electrons until they reach **octet** (or **duplet**) configurations of noble gases
  - **valence** – number of covalent bonds an atom forms (number of shared **e**$^-$ pairs)
  - the number of shared **e**$^-$ pairs equals the number of electrons an atom needs in order to complete its octet (or duplet) structure



- **Lewis structures** – diagrams showing the distribution of electrons in a molecule
  - **shared (bonding) e**$^-$ pairs – between the atoms (can be expressed as lines representing bonds)
  - **lone e**$^-$ pairs – not involved in bonding (not shared)

Example: Write the Lewis structure of HCl and determine the number of shared and lone **e**$^-$ pairs.



3 lone pairs at Cl and 1 bonding (shared) pair

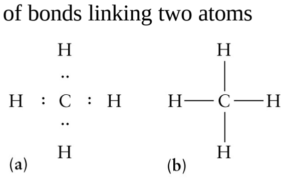
The Structures of Polyatomic Species

8.7 Lewis Structures

- Single bond – a single shared **e**$^-$ pair
- Multiple bonds – double or triple bonds (2 or 3 shared **e**$^-$ pairs)
- Bond order – number of bonds linking two atoms

Example: Methane, CH_4.

:Ċ· H· H· H· H·



3 Methane, CH_4

- Arrangement of atoms in polyatomic species (skeleton structure)
  - **central atom** – usually the atom with the lowest ***I*** (often written first in the formula) PCl_5, SO_3, …
  - normally **H** is not a central atom
  - normally the atoms are arranged symmetrically around the central atom CO_2 → OCO, OF_2 → FOF
- Polyatomic ions
  - cations and anions of an ionic compound are treated separately
  - total number of valence **e**$^-$ is adjusted for the charge of the ion

- Rules for writing Lewis structures
  1 write the skeleton structure by placing a single bond (**e**$^-$ pair) between each bonded pair of atoms
  2 count the total number of valence electrons, $\mathbf{n_{tot}}$, of all atoms (correct for the charges of ions)
  3 count the number of remaining electrons, $\mathbf{n_{rem}}$ (total number of **e**$^-$ minus **e**$^-$ used in the skeleton structure)
  4 count the number of needed electrons, $\mathbf{n_{need}}$ (the **e**$^-$ needed to complete the octet (or duplet) structures of all atoms)
  5 if $\mathbf{n_{need}} = \mathbf{n_{rem}}$, add the remaining **e**$^-$ as lone pairs to complete the octets for all atoms, or
  if $\mathbf{n_{need}} > \mathbf{n_{rem}}$ add multiple bonds (1 bond for each deficient pair of **e**$^-$) and complete the structure with lone pairs

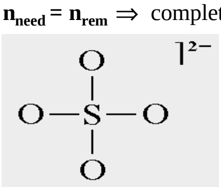
Example: Write the Lewis structure of **HCN**.

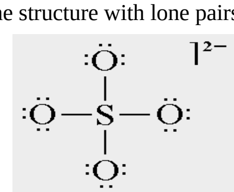
1. **C** is the central atom (lower ***I*** than **N**)
   ⇒ **H–C–N** (**4 e**$^-$ in the skeleton structure)
2. $\mathbf{n_{tot}}$ **= 1(H) + 4(C) + 5(N) = 10**
3. $\mathbf{n_{rem}}$ **= 10 - 4 = 6**
4. $\mathbf{n_{need}}$ **= 0(H) + 4(C) + 6(N) = 10**
5. $\mathbf{n_{need}} > \mathbf{n_{rem}}$ deficiency of **4 e**$^-$ (**2 e**$^-$ pairs)
   ⇒ add **2** more bonds between **C** and **N** and complete the structure with lone pairs

H–C≡N:

Example: Write the Lewis structure of $\mathbf{SO_4^{2-}}$.

1. **S** is the central atom (lower ***I*** than **O**)
2. **8 e**$^-$ in the skeleton structure
   $\mathbf{n_{tot}}$ **= 6(S) + 4×6(O) + 2(charge) = 32**
3. $\mathbf{n_{rem}}$ **= 32 - 8 = 24**
4. $\mathbf{n_{need}}$ **= 0(S) + 4×6(O) = 24**
5. $\mathbf{n_{need}} = \mathbf{n_{rem}}$ ⇒ complete the structure with lone pairs

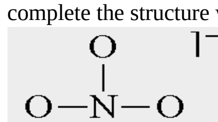


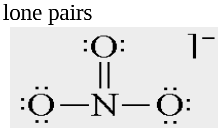


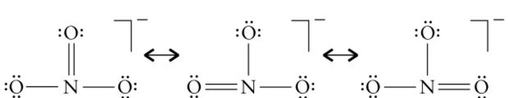
8.8 Resonance in Lewis Structures

Example: Write the Lewis structure of $\mathbf{NO_3^-}$.

1. **N** is the central atom (lower ***I*** than **O**)
2. **6 e**$^-$ in the skeleton structure
   $\mathbf{n_{tot}}$ **= 5(N) + 3×6(O) + 1(charge) = 24**
3. $\mathbf{n_{rem}}$ **= 24 - 6 = 18**
4. $\mathbf{n_{need}}$ **= 2(N) + 3×6(O) = 20**
5. $\mathbf{n_{need}} > \mathbf{n_{rem}}$ deficiency of **2 e**$^-$ (**1 e**$^-$ pair)
   ⇒ add **1** more bond between **N** and one of the **Os** and complete the structure with lone pairs



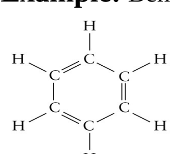




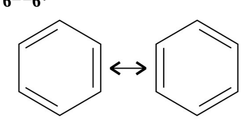
13 Nitrate ion, NO_3^-

- All three structures are valid Lewis structures and differ by the position of the **e**$^-$ pairs – **resonance structures**
- Neither of the resonance structures is realistic
- The real structure is a blend (resonance hybrid) of the contributing Lewis structures
- The three bonds are identical (intermediate between a single and a double bond)

Example: Benzene, $\mathbf{C_6H_6}$.

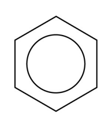


16 Kekulé structure



19 Benzene resonance structure

- The two structures have **equal energies** and **contribute equally** to the hybrid structure
- The electrons from the double bonds are **delocalized** between all six **C** atoms
- The six bonds are equivalent to each other



20 Benzene, C_6H_6