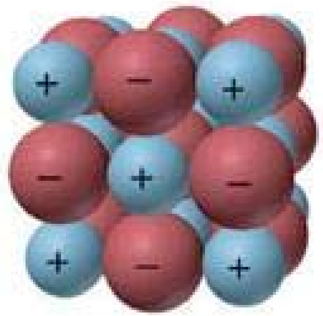
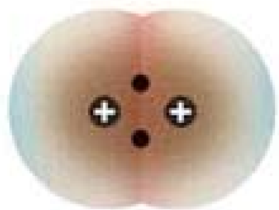
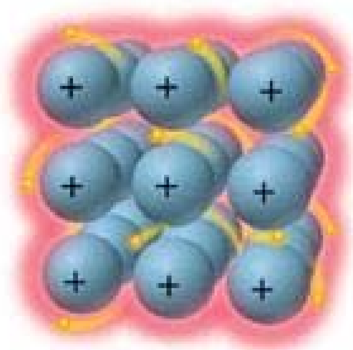
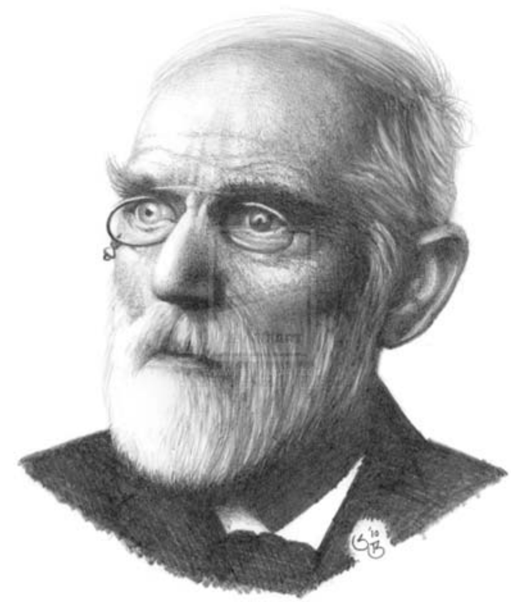
Force	Model	Basis of Attraction	Energy (kJ/mol)	Example
Bonding				
Ionic		Cation–anion	400–4000	NaCl
Covalent		Nuclei–shared e^- pair	150–1100	H—H
Metallic		Cations–delocalized electrons	75–1000	Fe

van der Waals interactions or Force

In physical chemistry, the van der Waals force (or van der Waals' interaction), named after Dutch scientist Johannes Diderik van der Waals, is the sum of the attractive or repulsive forces between molecules (or between parts of the same molecule) other than those due to covalent bonds, or the electrostatic interaction of ions with one another, with neutral molecules, or with charged molecules. The term includes:

- force between two permanent dipoles (Keesom force)
- force between a permanent dipole and a corresponding induced dipole (Debye force)
- force between two instantaneously induced dipoles (London dispersion force).

It is also sometimes used loosely as a synonym for the totality of ***intermolecular forces.***



Johannes Diderik van der Waals
Leiden, The Netherlands, 1837 -1923

Group 18 Elements

Element	*mp, °C*	*bp, °C*
He	−272	−269
Ne	−249	−246
Ar	−189	−186
Kr	−157	−152
Xe	−112	−107
Rn	−71	−62

Molecular Halogens

Compound	*mp, °C*	*bp, °C*
F_2	−220	−188
Cl_2	−101	−35
Br_2	−7	59
I_2	114	184

Silicon Tetrahalides

Compound	*mp, °C*	*bp, °C*
SiF_4	−90	−86
$SiCl_4$	−70	58
$SiBr_4$	5	154
SiI_4	120	288

Non-Polar:
MPt or BPt increases with increasing mass or volume

London Dispersion Forces

Chemical Forces and Interactions: Dependence on Distance between Centers

Table 6.3 Summary of chemical forces and interactions

Type of interaction	Strength	Energy–distance function
Covalent bond	Very strong	Complex, but comparatively long-range
Ionic bond	Very strong	$1/r$, comparatively long-range
Ion–dipole	Strong	$1/r^2$, short-range
Dipole–dipole	Moderately strong	$1/r^3$, short-range
Ion-induced dipole	Weak	$1/r^4$, very short-range
Dipole-induced dipole	Very weak	$1/r^6$, extremely short-range
London dispersion energy	Very weak[a]	$1/r^6$, extremely short-range

[a] Since London forces increase with increasing size and there is no limit to the size of molecules, these forces can become rather large. In general, however, they are very weak.

Ion-Ion

$$E = \frac{Z^+Z^-}{4\pi r \varepsilon_0}$$

Ion-Dipole

$$E = -\frac{|Z^\pm|\,\mu}{4\pi r^2 \varepsilon_0}$$

Dipole-Dipole

$$E = \frac{-2\,\mu_1\mu_2}{4\pi r^3 \varepsilon_0}$$

Ion-Induced Dipole

$$E = -\frac{1}{2}\,\frac{Z^2\alpha}{r^4}$$

Dipole-Induced Dipole

$$E = \frac{-\mu_2\alpha}{r^6}$$

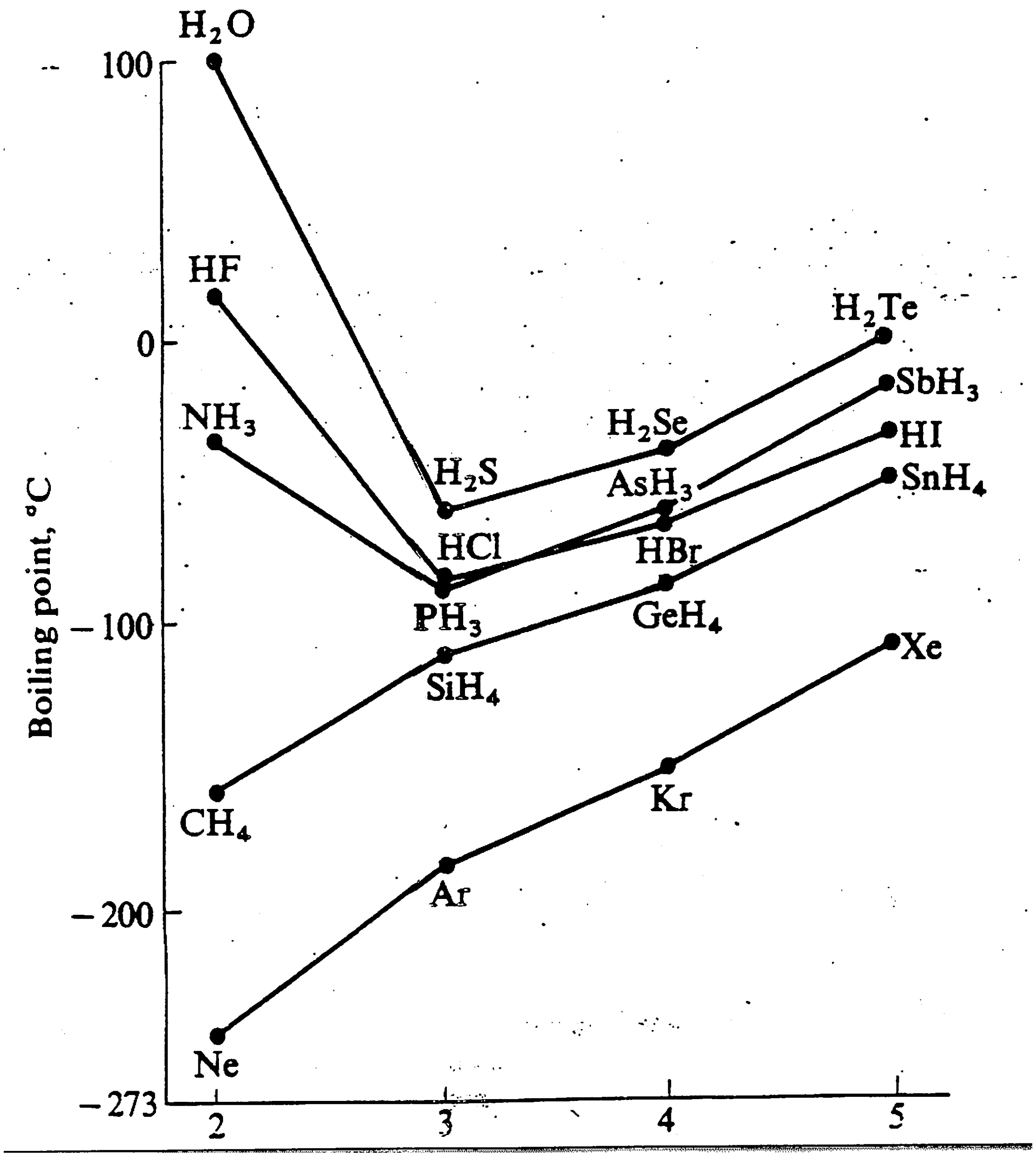
Instantaneous Dipole-Induced Dipole (London Dispersion Forces)

$$E = \frac{-2\bar{\mu}\alpha}{r^6}$$

$$E = \frac{-3I\alpha^2}{4r^6}$$

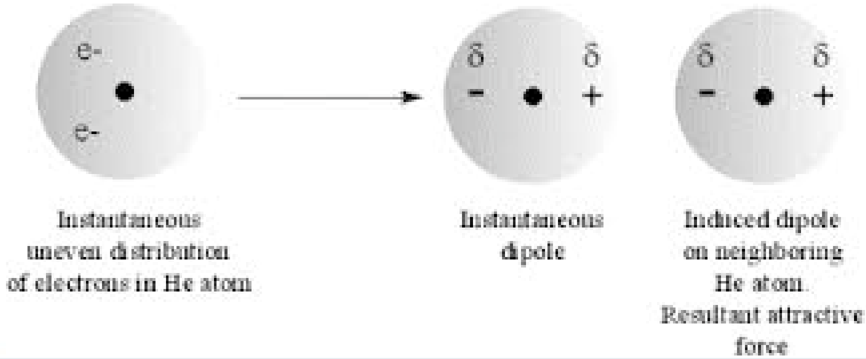
Intermolecular Forces

An Iconic graphic: Plots of boiling points of EH_x vs. Molecular weights



6.28

Instantaneous dipole/Induced dipole



Element	Melting point / K	$\Delta_{fus}H$(mp) / kJ mol^{-1}	Boiling point / K	$\Delta_{vap}H$(bp) / kJ mol^{-1}	van der Waals radius (r_v) / pm
Helium	‡	–	4.2	0.08	99
Neon	24.5	0.34	27	1.71	160
Argon	84	1.12	87	6.43	191
Krypton	116	1.37	120	9.08	197
Xenon	161	1.81	165	12.62	214
Radon	202	–	211	18	–

‡Helium cannot be solidified under atmospheric pressure, the pressure condition for which all other phase changes in the table are considered.

Vander Waal Forces of London type: Depend on molecular volume

Table 6.1 Selected physical data for group 18 elements.

Hydrogen bonding

DESCRIPTION (= DEFINITION) A hydrogen bond is the weak force of electrostatic attraction

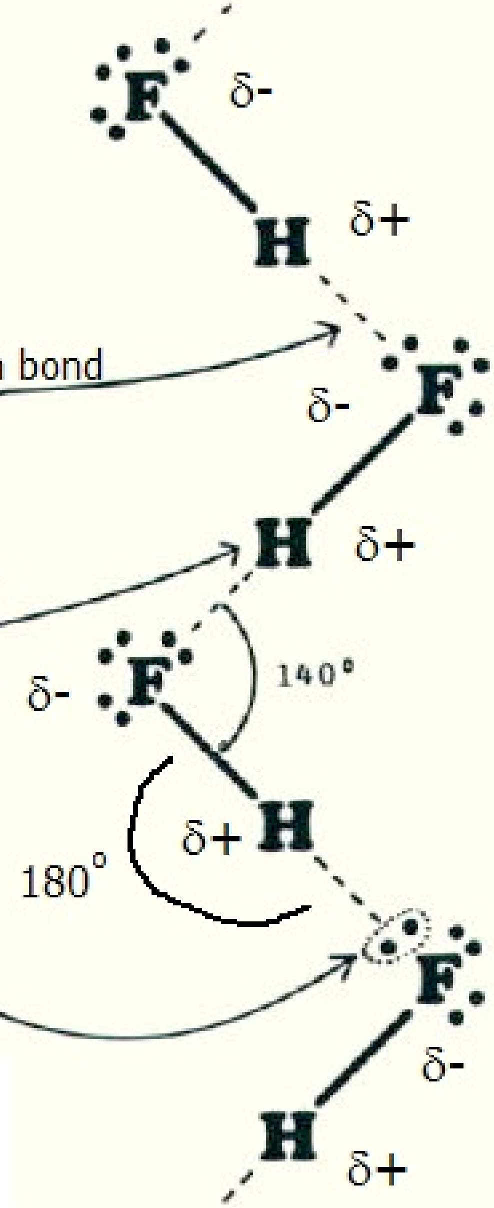
BETWEEN

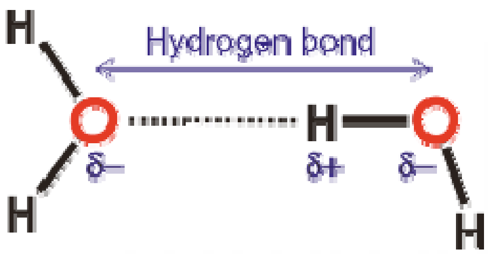
a partially positive hydrogen atom covalently bound to one of the three most electronegative elements (N, O or F)

AND

the lone pair on a covalently bound N, O or F atom.

Note that the hydrogen bond is in line with the H-F covalent bond. i.e. this bond angle is 180°





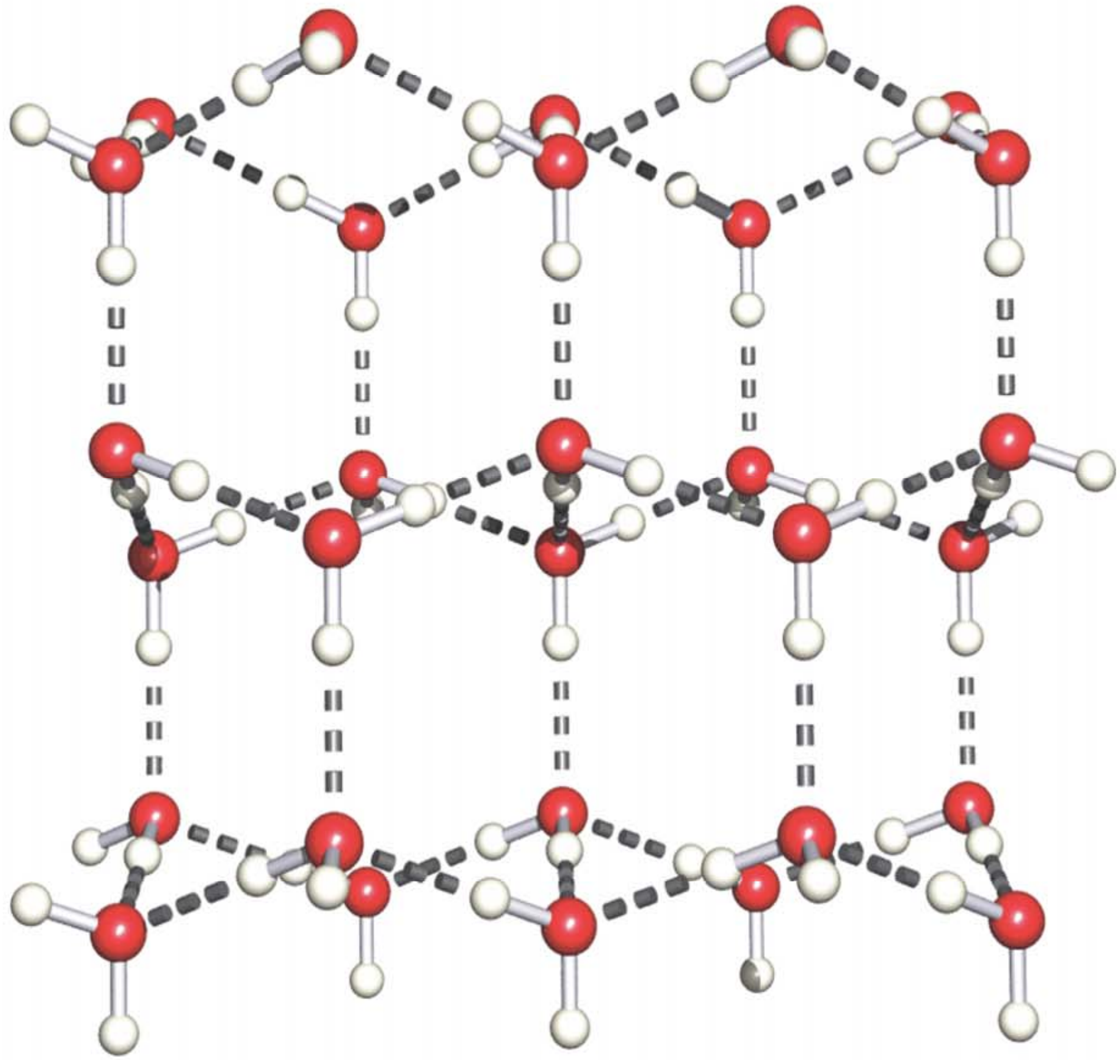
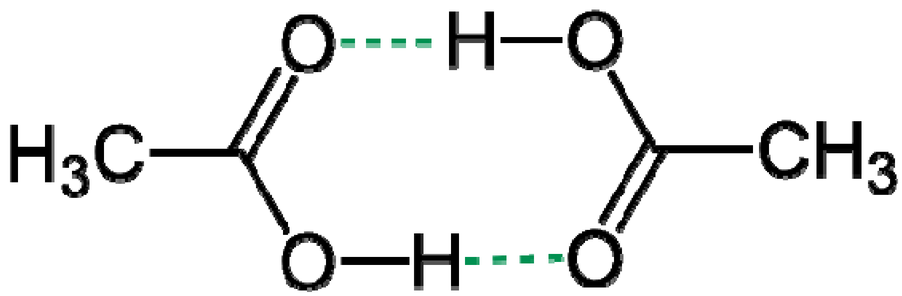
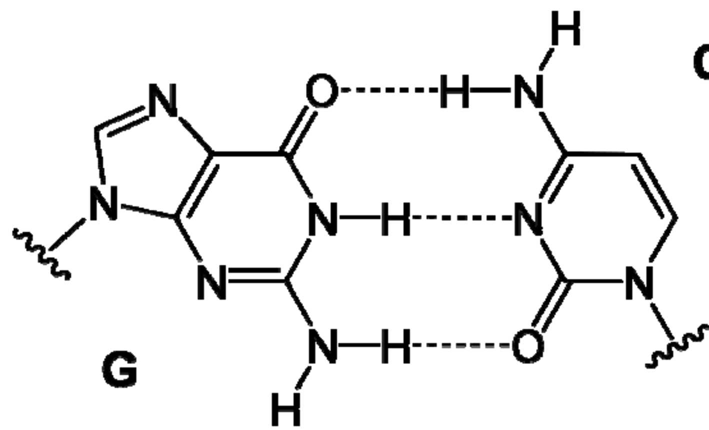


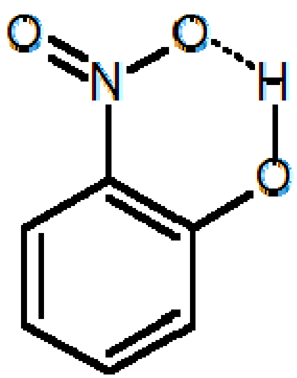
Fig. 7.1 Part of the structure of ordinary ice; it consists of a three-dimensional network of hydrogen-bonded H_2O molecules.

Intermolecular H-bonding



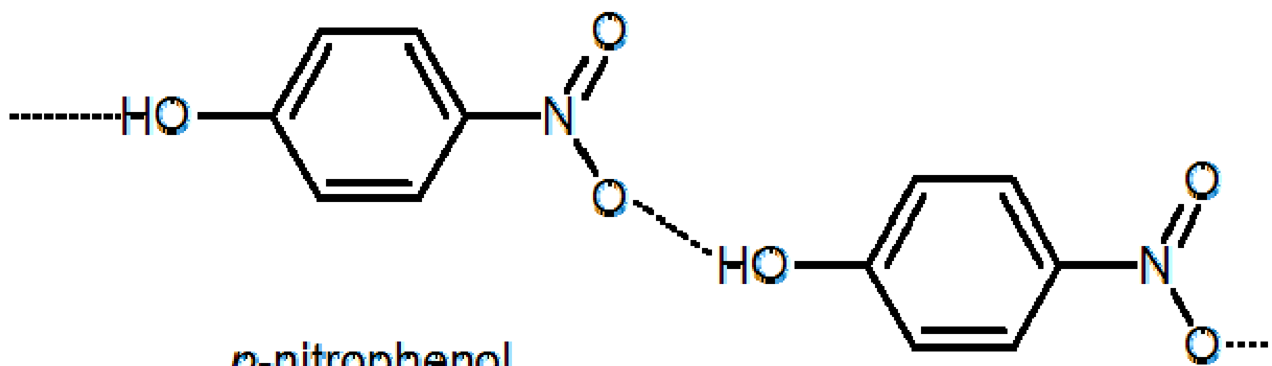


Intramolecular H-bonding



o-nitrophenol

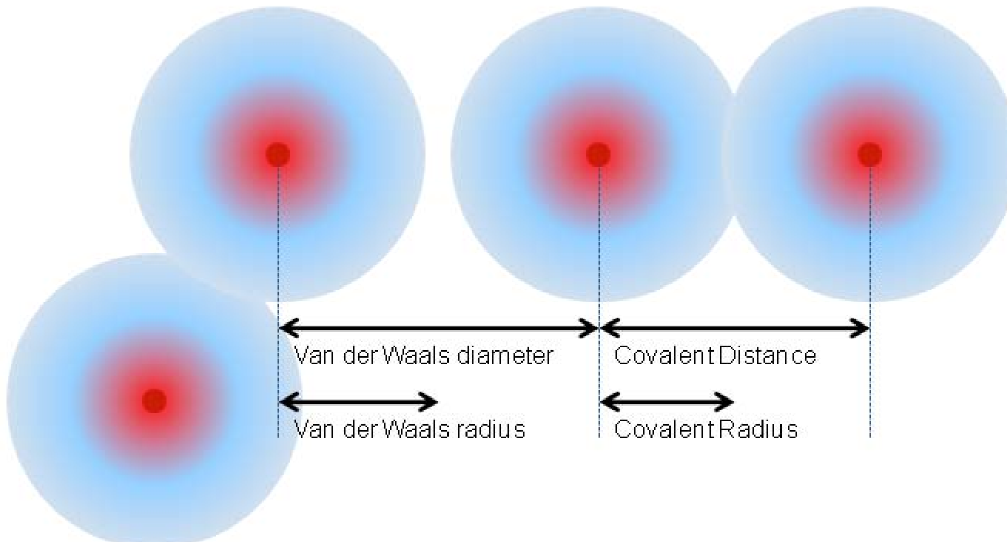
(intramolecular hydrogen bonding)



p-nitrophenol

(intermolecular hydrogen bonding)

Van der Waals radius (non-bonding distance)



Hydrogen bond strengths:

F−H···:F (161.5 kJ/mol or 38.6 kcal/mol), illustrated uniquely by HF_2^-, bifluoride

O−H···:N (29 kJ/mol or 6.9 kcal/mol), illustrated water-ammonia

O−H···:O (21 kJ/mol or 5.0 kcal/mol), illustrated water-water, alcohol-alcohol

N−H···:N (13 kJ/mol or 3.1 kcal/mol), illustrated by ammonia-ammonia

N−H···:O (8 kJ/mol or 1.9 kcal/mol), illustrated water-amide

HO−H···:OH_3^+ (18 kJ/mol[19] or 4.3 kcal/mol)

Is there a H-bond possible in $CHCl_3$?