

CHEMICAL BONDS

Lecture 5



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RECAP

- electrons have both **POTENTIAL** and **KINETIC** energy
- the chemical energy is a combination of kinetic and potential energy (total energy, $E = E_P + E_K$)
- **PE** of the outermost electrons depends on the position of the electron in the atom (electron configuration)
- **periodic table** - useful tool to derive properties of the atom (reactivity, bonds it can form, etc.)

TODAY'S LECTURE

- how much energy is stored in chemical bonds?
- types of bonds between atoms to consider:
 - ionic
 - metallic
 - covalent (σ , π)



OCTET / DUET RULE

refers to the tendency of atoms to have two or eight electrons in the valence shell

1 s^1	2 s^2	3 s^2p^1	4 s^2p^2	5 s^2p^3	6 s^2p^4	7 s^2p^5	8/0 s^2p^6
H 1							He 2
Li 3	Be 4	B 5	C 6	N 7	O 8	F 9	Ne 10
Na 11	Mg 12	Al 13	Si 14	P 15	S 16	Cl 17	Ar 18
K 19	Ca 20	Ga 31	Ge 32	As 33	Se 34	Br 35	Kr 36
Rb 37	Sr 38	In 49	Sn 50	Sb 51	Te 52	I 53	Xe 54
Cs 55	Ba 56	Tl 81	Pb 82	Bi 83	Po 84	At 85	Rn 86
Fr 87	Ra 88	Nh 113	Fl 114	Mc 115	Lv 116	Ts 117	Og 118

Metal atoms want 0 valence electrons

Non-metal/Metalloid atoms want 8 valence electrons (H wants 2)

Metal + Non-metal/Metalloid

electron transfer
ionic compound

Non-metal/Metalloid + Non-metal/
Metalloid

sharing electrons
covalent compound



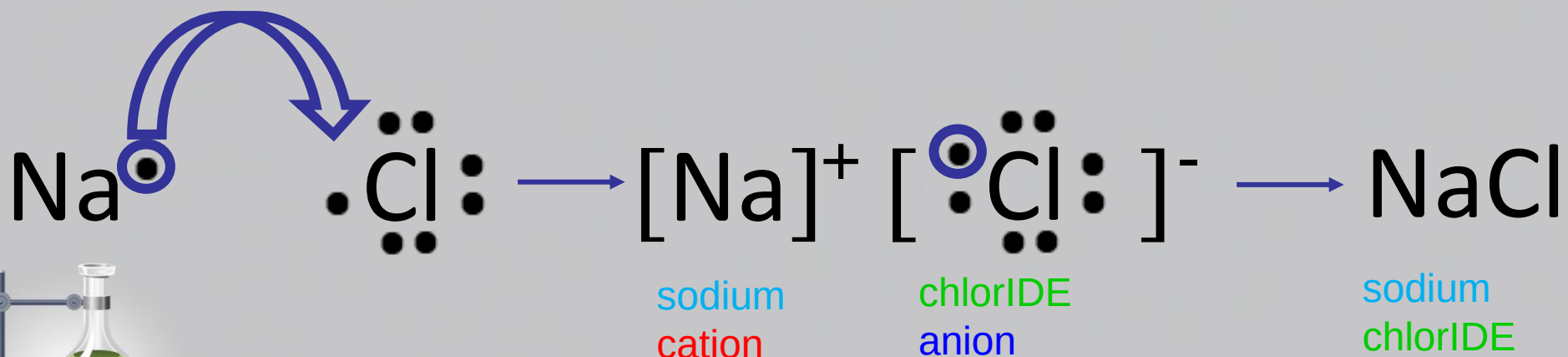
IONIC BONDS

- exists between oppositely charged ions (atoms or groups of atoms with an electrostatic charge; **ANION**⁻, **CATION**⁺) that electrostatically attract each other
- forms when electrons are transferred from a **metal** to a **non-metal** in order to obtain a full valence shell for both atoms (8 electrons-octet)

EXAMPLE, NaCl

Na $1s^2 \ 2s^2 \ 2p^6 \ 3s^1$

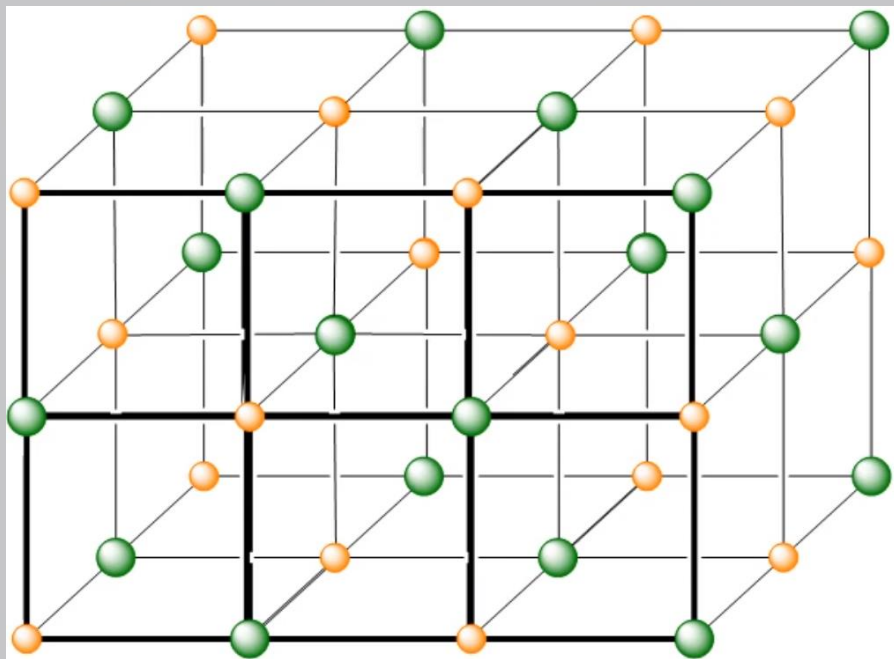
Cl $1s^2 \ 2s^2 \ 2p^6 \ 3s^2 \ 3p^5$



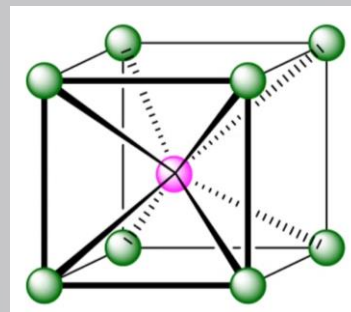
LATTICE ENERGY

An ionic compound is stable because of the electrostatic attraction between its positive and negative ions. The **lattice energy** of a compound is a measure of the strength of this attraction.

SODIUM CHLORIDE



CAESIUM CHLORIDE



$$PE = q_1 q_2 / 4\pi\epsilon_0 r$$

q_1 and q_2 separated by r
(charge on electron $-1,602 \times 10^{-19}$ C)

ϵ_0 - permittivity of a medium

range of bond energies 600 – 12,000 kJ/mol

VERY, VERY STRONG



IONIC COMPOUNDS IN OME

explosive compositions,
propellants:

nitrates:

potassium nitrate



ammonium nitrate



chlorates:

potassium chlorate



perchlorates:

potassium perchlorate



pyrotechnics

			
RED STRONTIUM SALTS Strontium Nitrate Strontium Carbonate Strontium Sulfate	ORANGE CALCIUM SALTS Calcium Carbonate Calcium Chloride Calcium Sulfate	YELLOW SODIUM SALTS Sodium Nitrate Sodium Oxalate Cryolite	GREEN BARIUM SALTS Barium Nitrate Barium Carbonate Barium Chloride Barium Chlorate
			
BLUE COPPER SALTS Copper (I) Chloride Copper Carbonate Copper Oxide	PURPLE COMBINE Copper & Strontium Compounds	SILVER WHITE HOT Magnesium & Aluminum	WHITE BURNING METAL Magnesium Aluminum Titanium

COVALENT BONDS

- involves the sharing of **electron pairs** between atoms
- for many molecules, the sharing of electrons allows each atom to attain the equivalent of a full outer shell (**octet**), corresponding to a stable electronic configuration.

In organic chemistry, covalent bonds are much more common than ionic bonds.

- covalent bonding includes many kinds of interactions, including σ -bonding, π -bonding, metal-to-metal bonding.

range of bond energies 200 – 1,000 kJ/mol



COVALENT BONDS - LEWIS STRUCTURES

simplified representation of the valence shell electrons in a molecule

- Symbol and valence electrons (as dots)
- Top, bottom, left right – no set order
- Single until 4 of them, then pair up

Sodium atom	1 v.e.	Na •	Sodium cation	0 v.e.	[Na] ⁺
Carbon atom	4 v.e.	• • C • •			
Nitrogen atom	5 v.e.	• • N • ••			
Chlorine atom	7 v.e.	•• • Cl • ••	Chloride anion	8 v.e.	[: Cl :] ⁻

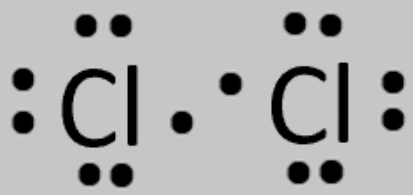


As sodium and chlorine atoms tend to form ionic compounds, there are shown here for demonstration of concept

COVALENT MOLECULES

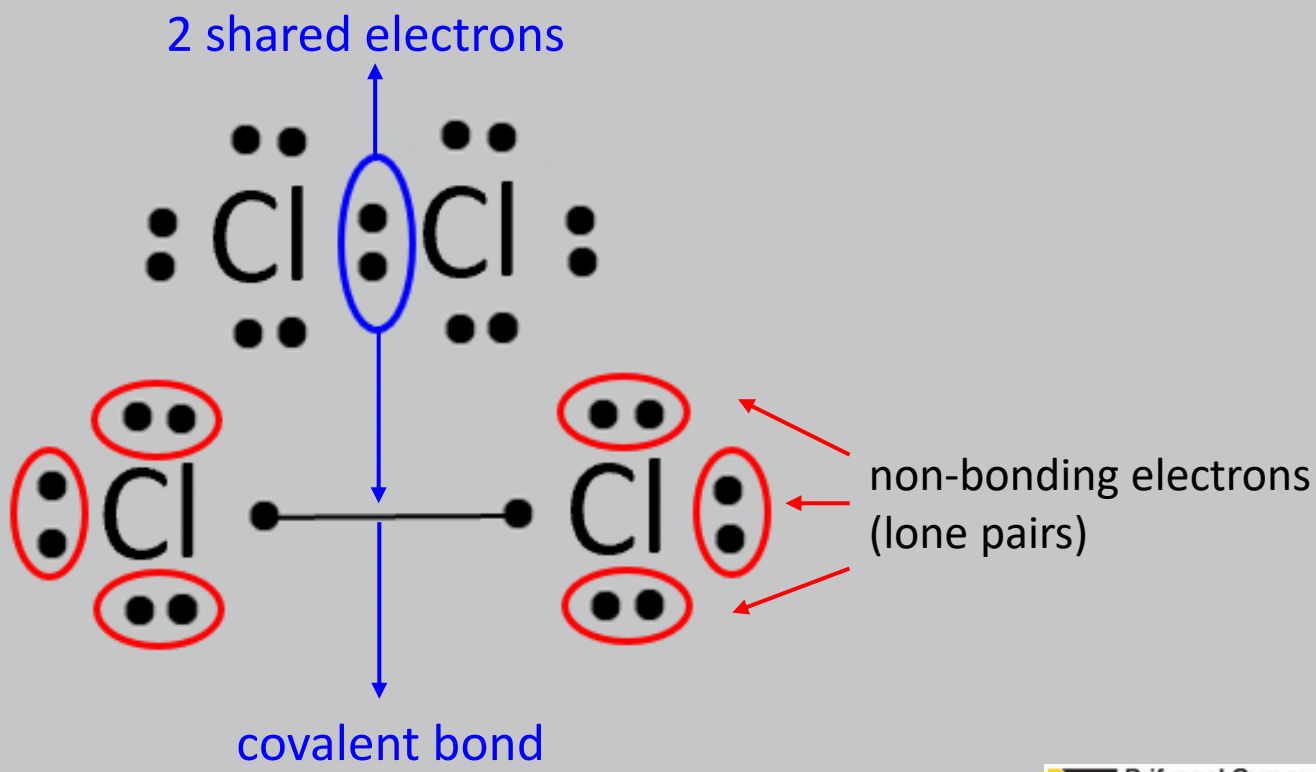


7 valence electrons
needs **one** more

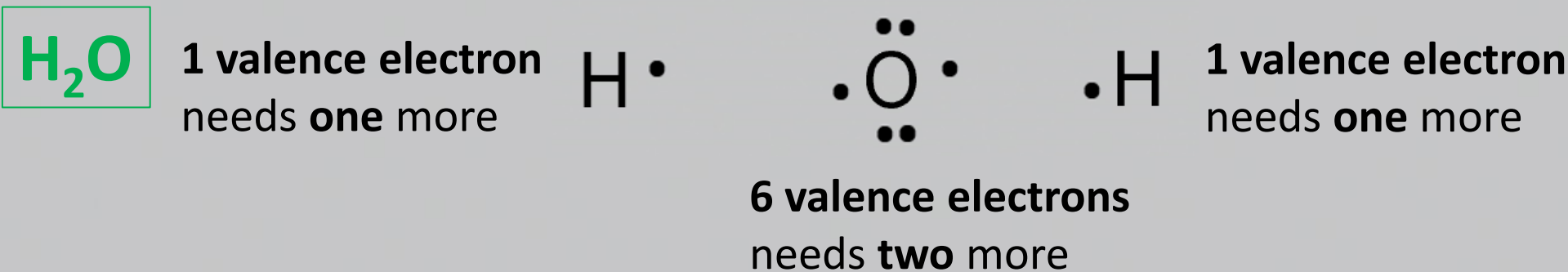


7 valence electrons
needs **one** more

Chlorine atom wants 1 electron to make one bond



COVALENT MOLECULES



Hydrogen atom wants 1 electron to make one bond

Oxygen atom wants 2 electrons to make two bonds

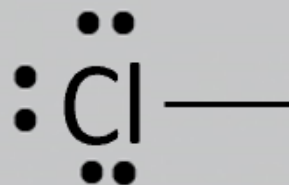


COVALENT MOLECULES

If an atom wants **X more electrons**, it will make **X bonds** (if sensible)

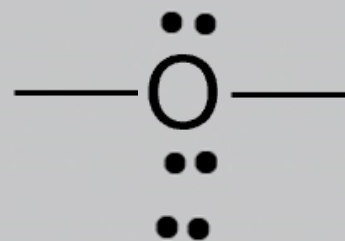
Cl 7 v.e.

wants 1 more to make 1 bond



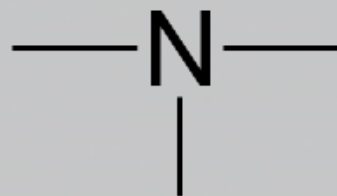
O 6 v.e.

wants 2 more to make 2 bonds



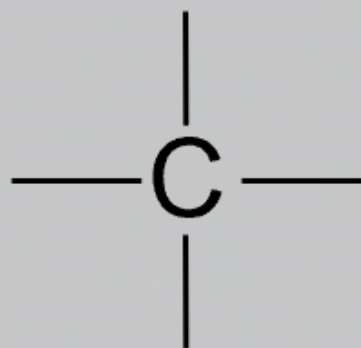
N 5 v.e.

wants 3 more to make 3 bonds



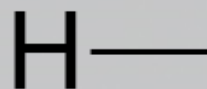
C 4 v.e.

wants 4 more to make 4 bonds



H 1 v.e.

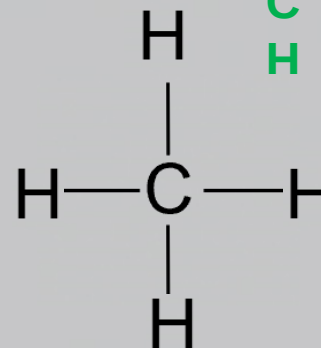
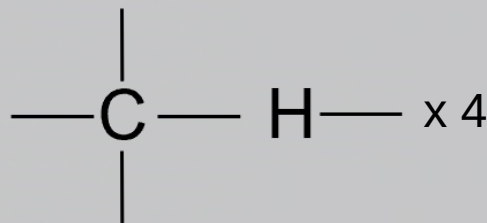
wants 1 more to make 1 bond



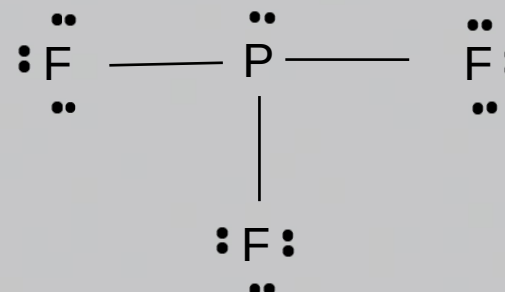
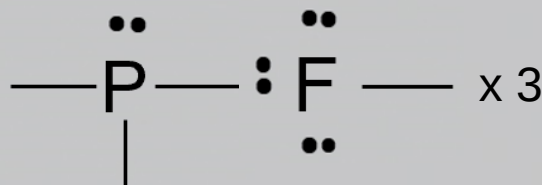
COVALENT MOLECULES: single bonds

F	etc	makes 1 bond
O	etc	makes 2 bonds
N	etc	makes 3 bonds
C	etc	makes 4 bonds
H		makes 1 bond

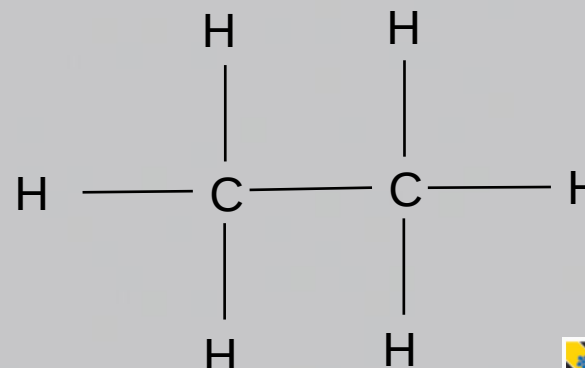
CH₄



PF₃

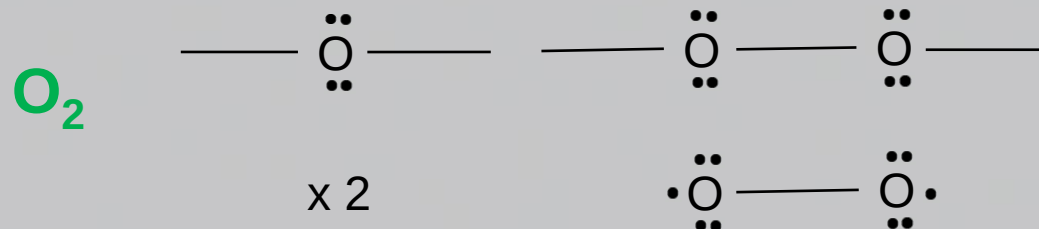


C₂H₆

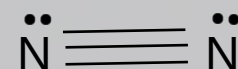
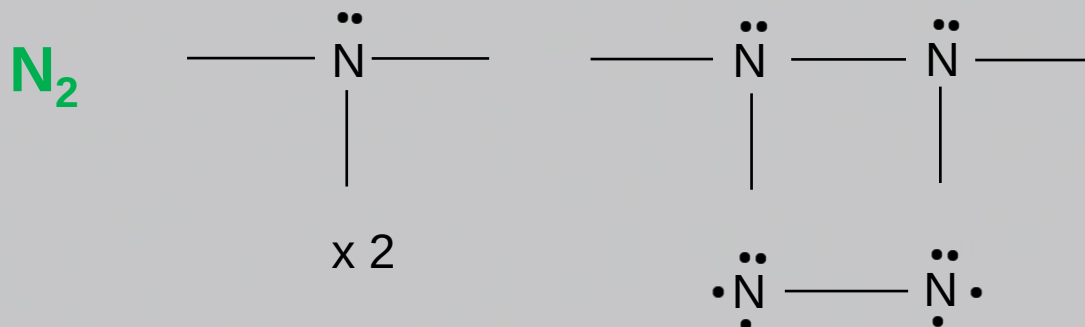


COVALENT MOLECULES: multiple bonds

F etc	makes 1 bond
O etc	makes 2 bonds
N etc	makes 3 bonds
C etc	makes 4 bonds
H	makes 1 bond



4 shared electrons
DOUBLE BOND



6 shared electrons
TRIPLE BOND

very important in the chemistry of explosives



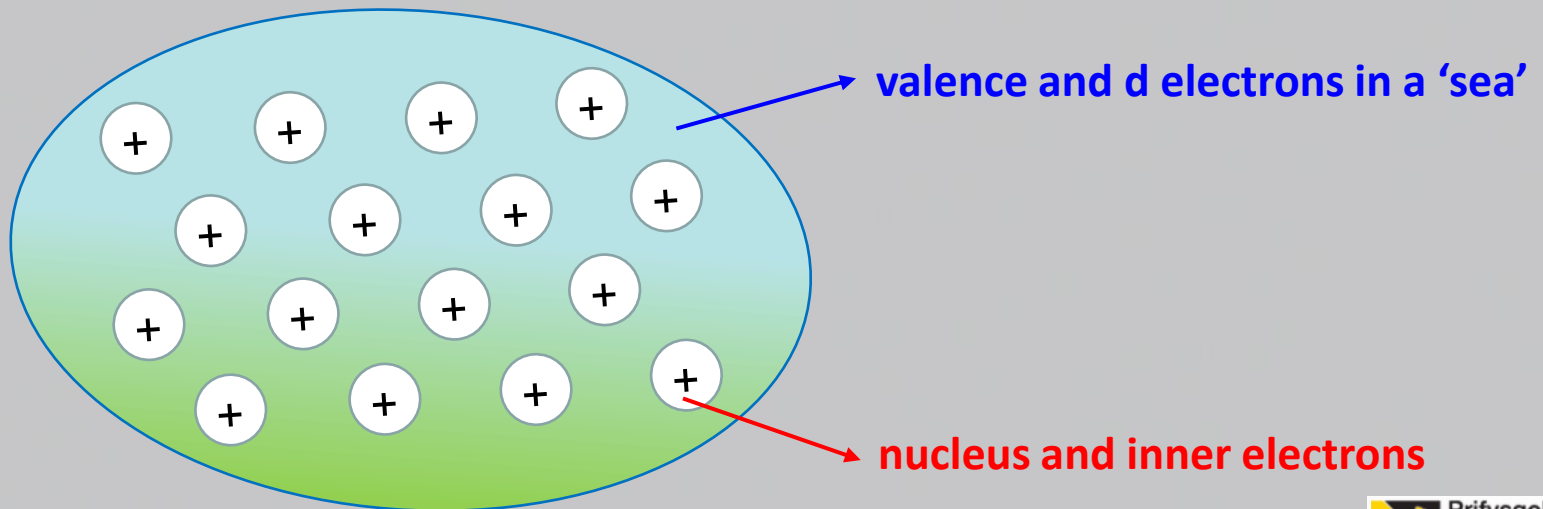
QUESTION 1: How many lone pairs will be present on a nitrogen, sulphur, carbon atom when it obeys 'common sense' (i.e. achieves 8 valence electrons by making regular covalent bond(s))?

QUESTION 2: Draw the Lewis structure of NH_3 , C_2F_6 . How many lone pairs are there?



METALLIC BOND

- **metallic bond** - electrostatic attractive force between conduction electrons that holds closely packed atoms together in a metallic substance
- the **valence electron** shell of each atom overlaps with a large number of neighboring atoms so the **valence electrons** move from one atom to another and are not associated with any specific pair of atoms
- **valence electrons** are nonlocalized, moving freely throughout the entire crystal
- the atoms that the **electrons** leave behind become positive ions, and the interaction between such ions and valence electrons gives rise to the binding force that holds the metallic crystal together



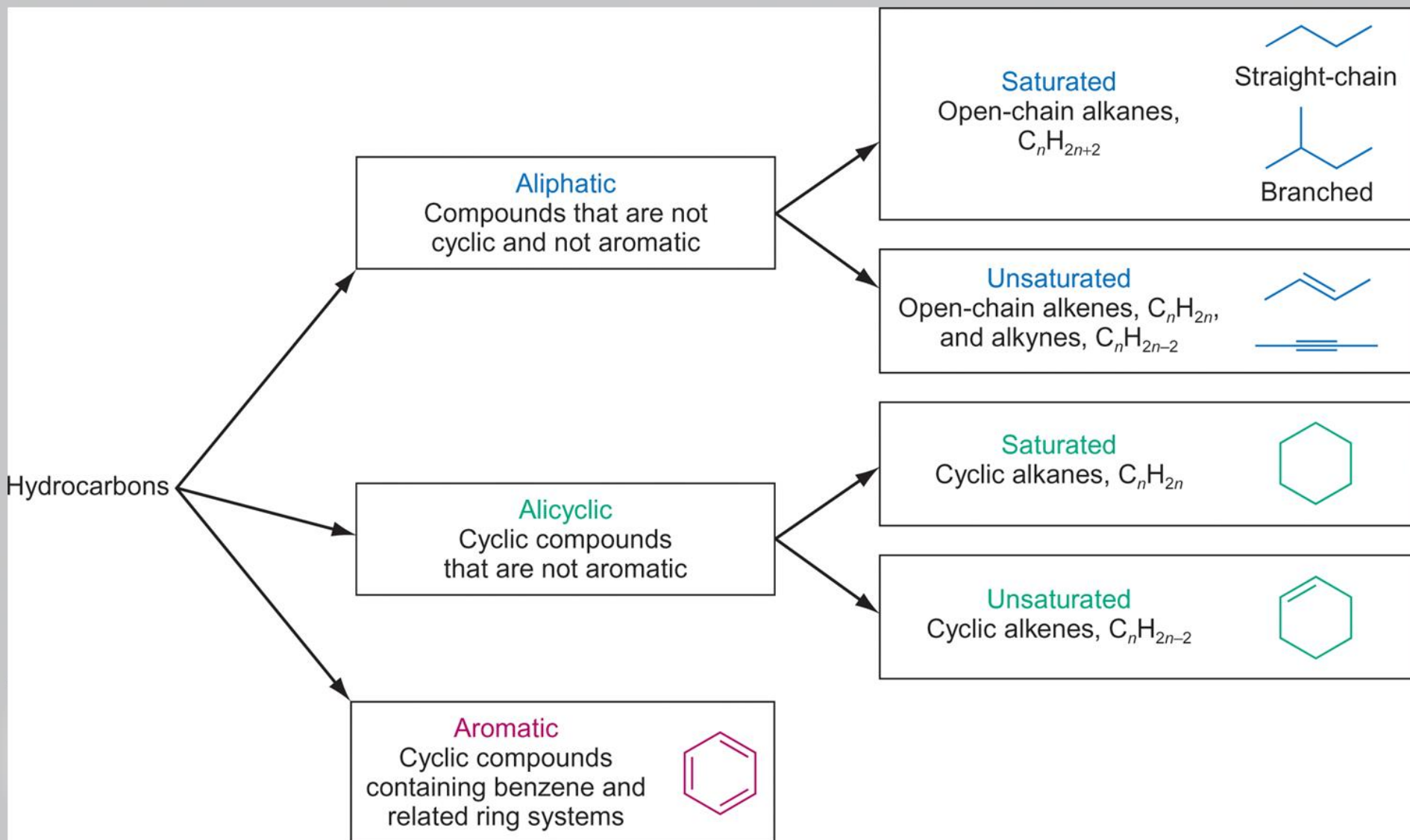
STRONG TO VERY, VERY STRONG



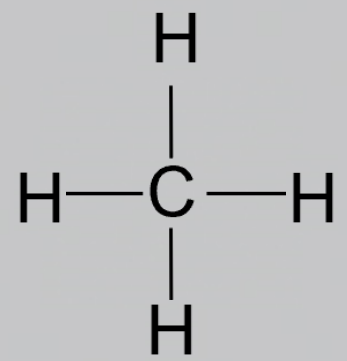
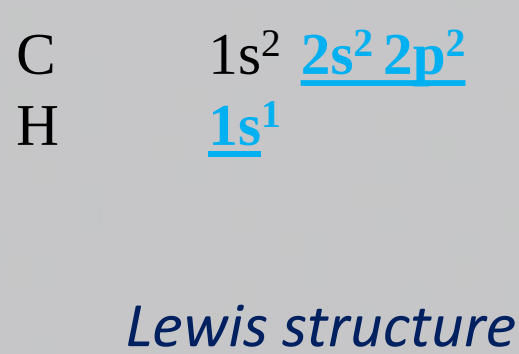
HYDROCARBON COMPOUNDS

most HE compounds are the derivatives of hydrocarbons

(an organic compound consisting entirely of hydrogen and carbon)



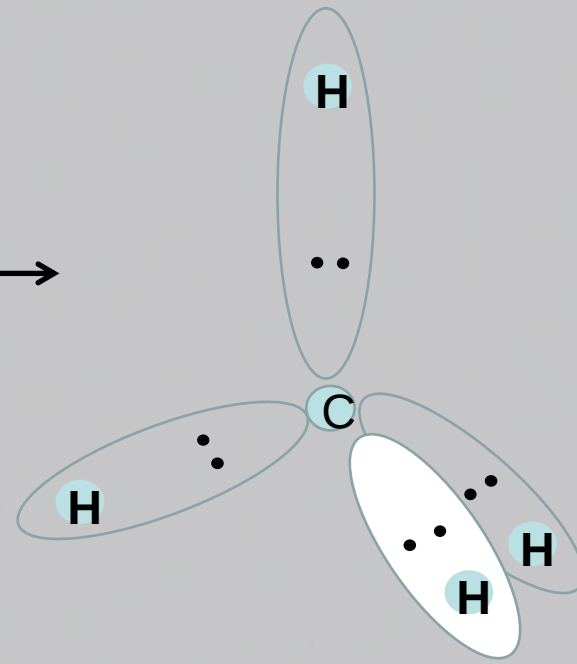
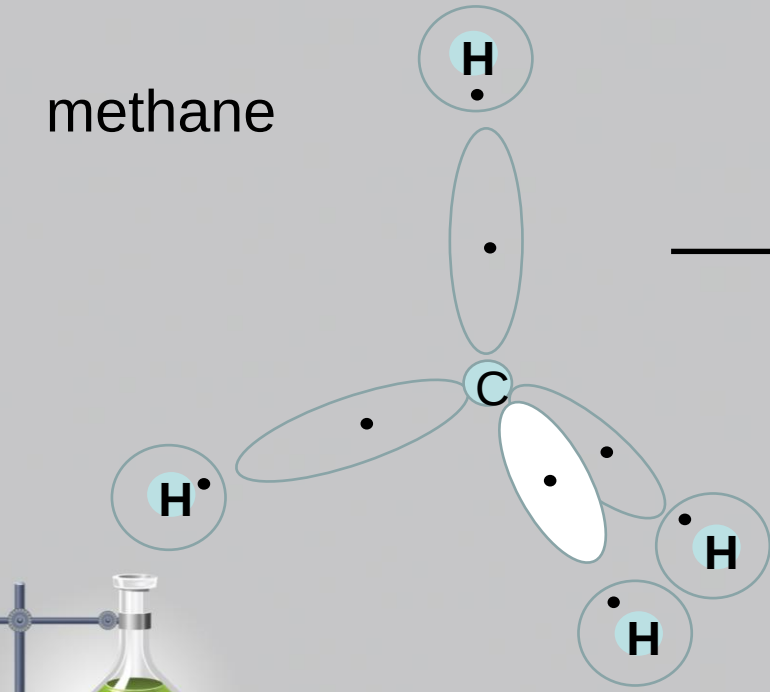
ALIPHATIC HYDROCARBON STRUCTURES – CH₄



TETRAHEDRON
angle 109.5°

molecular structure

methane

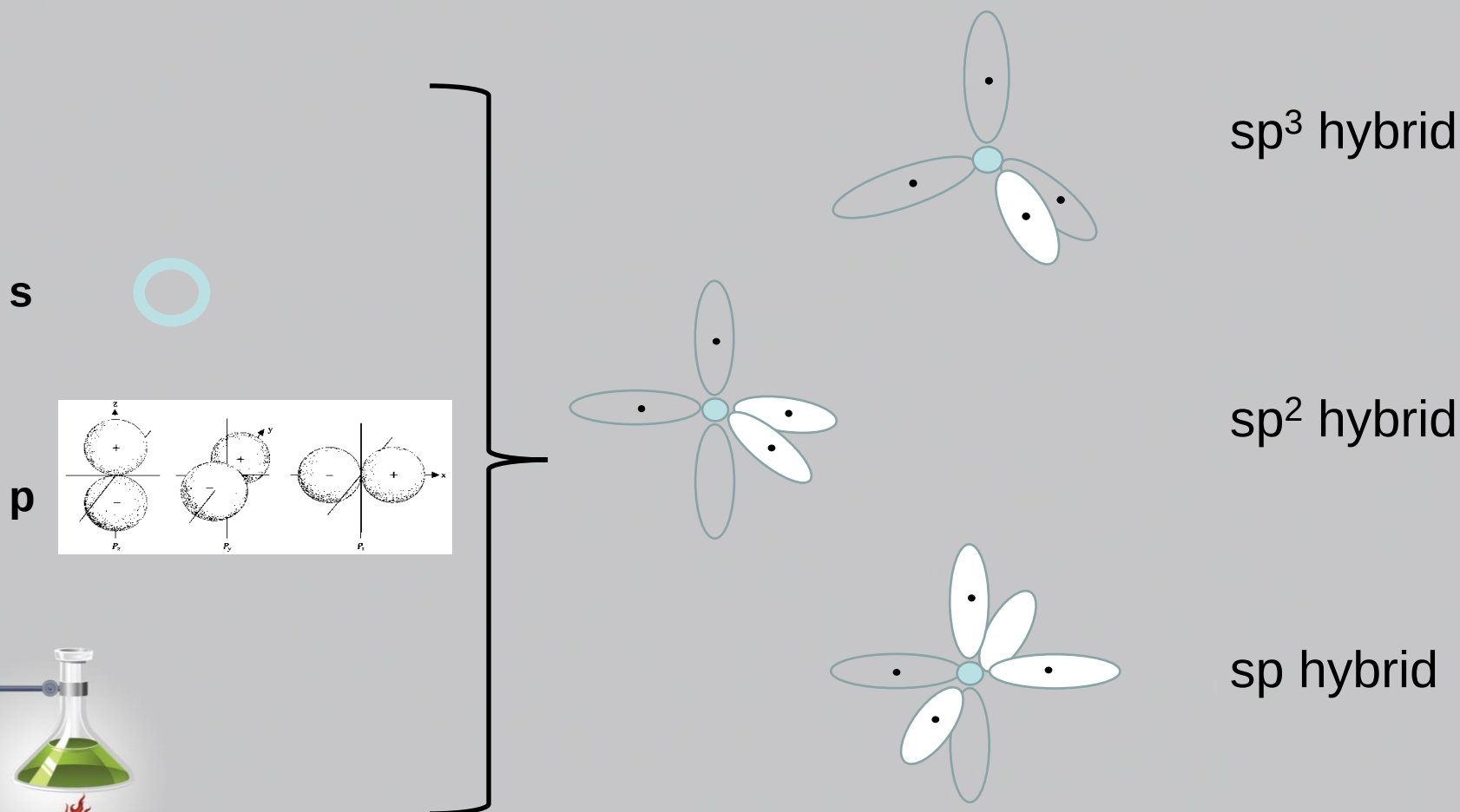


4 single bonds from
mixed orbitals s and p



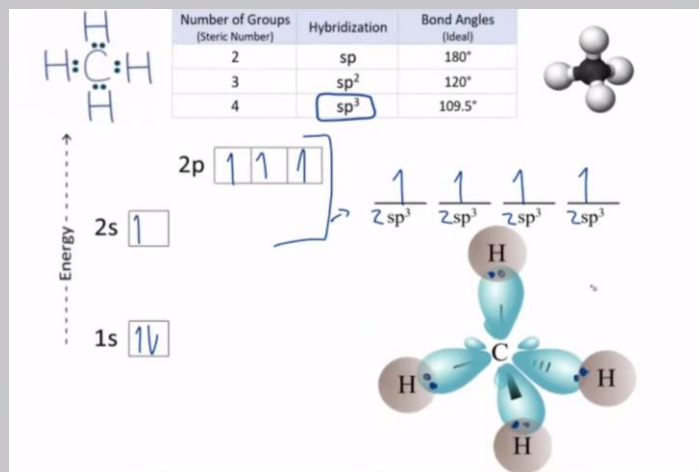
COVALENT BONDS – HYBRID ORBITALS

The 2s and 2p orbitals of each carbon atom undergo a transformation to form *hybrid orbitals* which bond with hybrids of other carbon atoms or with the 1s orbital of a hydrogen atom.

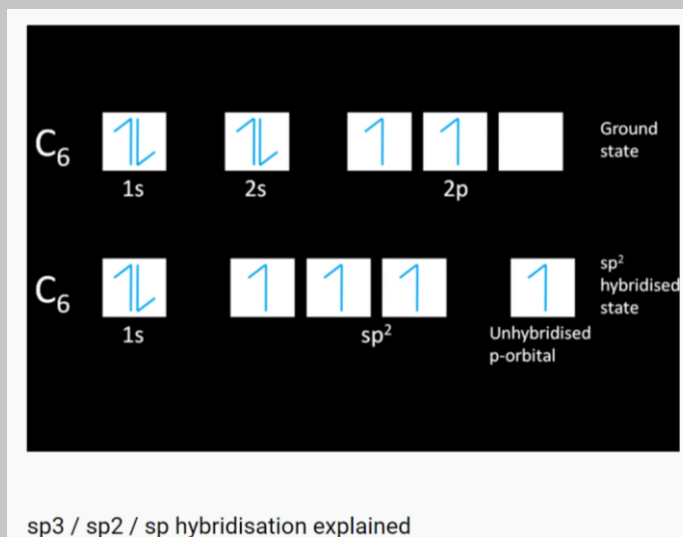


HYBRIDIZATION VIDEOS

<https://youtu.be/Pw3RUzQiFWo>

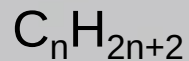


<https://youtu.be/5HvowNWIg5E>



ALIPHATIC HYDROCARBON - SATURATED

Alk**ane**s



Electronic structures

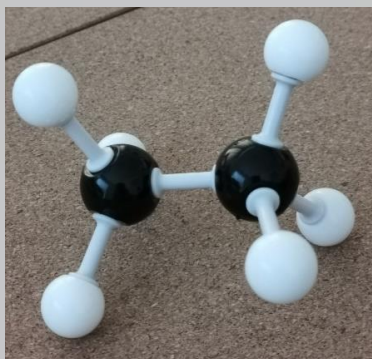
H $1s^1$

C $1s^2 2s^2 2p^2$

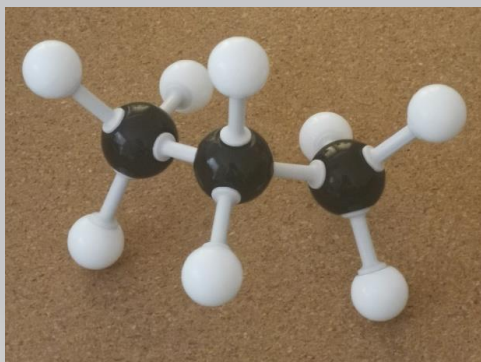
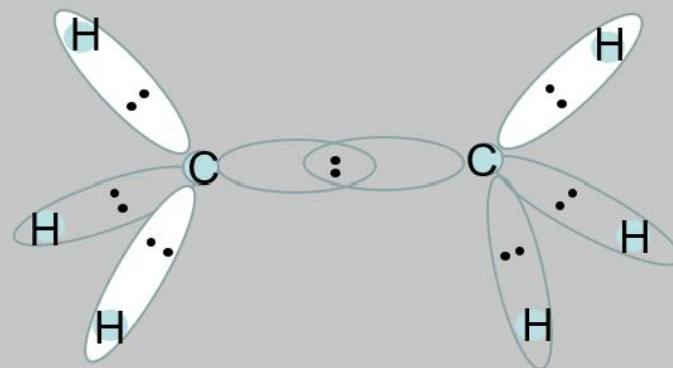
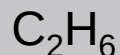
single covalent bonds, sp^3 hybridization $\rightarrow$ σ bonds



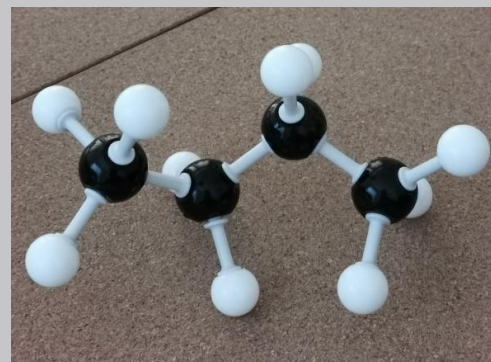
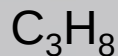
meth**ane**



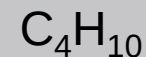
eth**ane**



prop**ane**

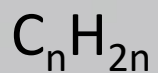


but**ane**

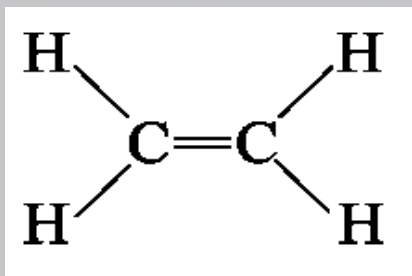
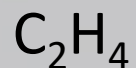


ALIPHATIC HYDROCARBON - UNSATURATED

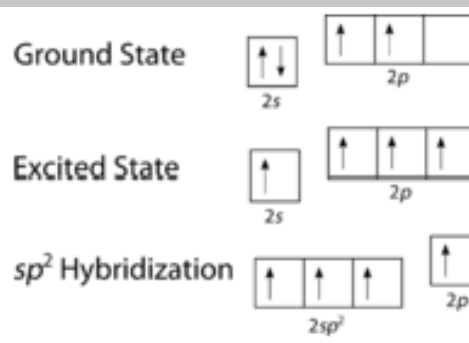
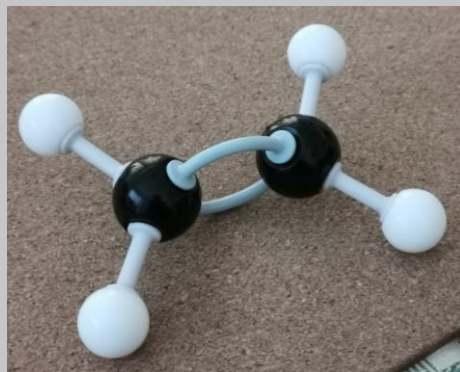
Alkenes



ethene



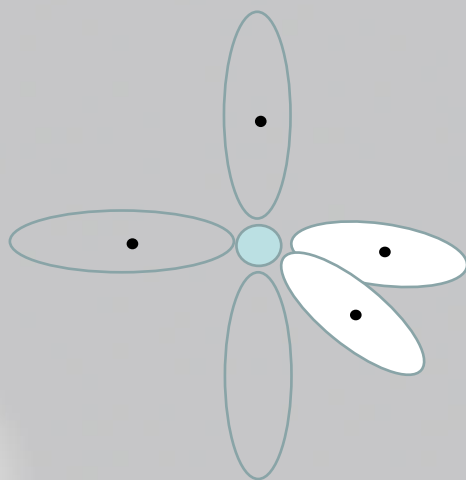
Lewis structure



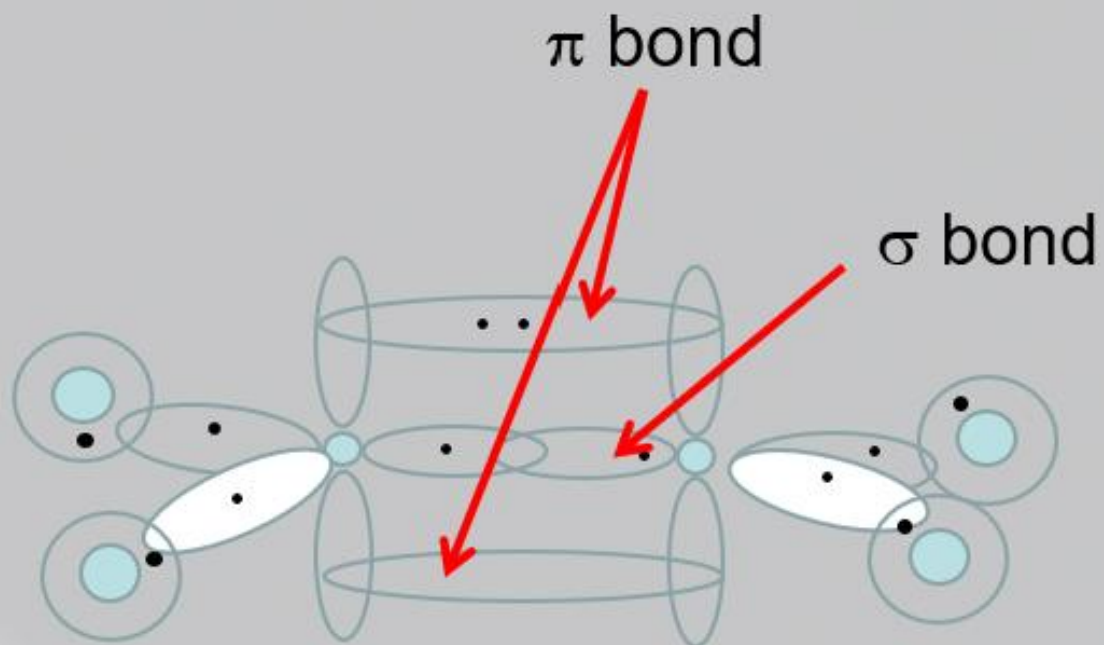
TRIGONAL PLANAR

angle 120°

molecular structure

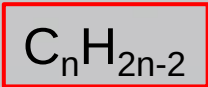


sp^2 hybrid orbital

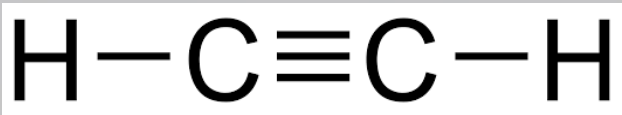


ALIPHATIC HYDROCARBON - UNSATURATED

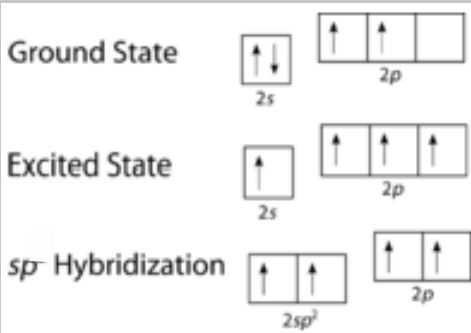
Alkynes



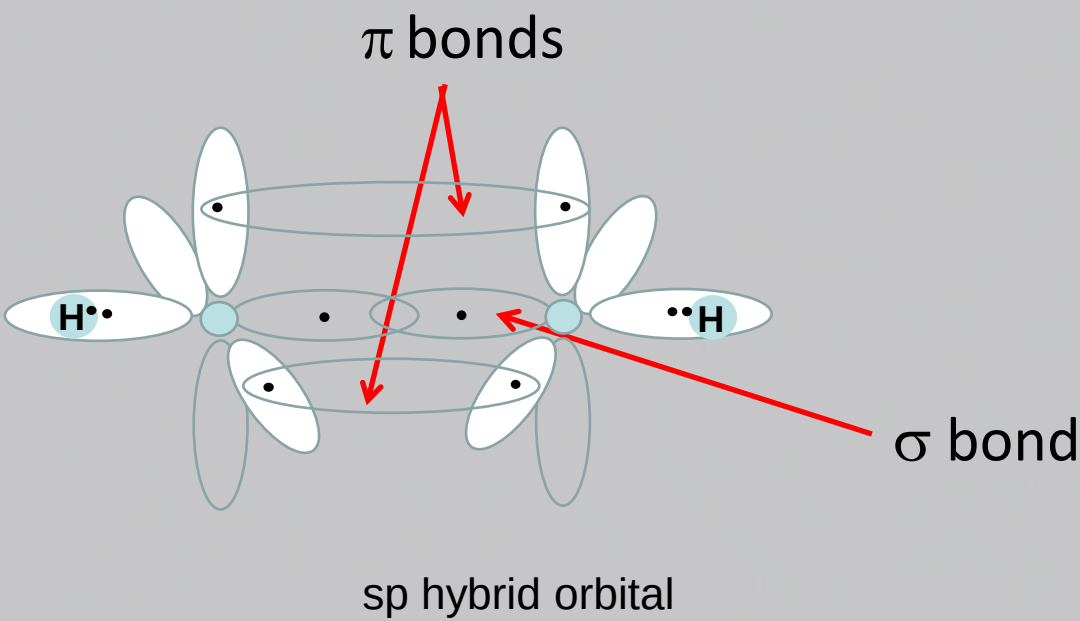
ethyne (acetylene)
 C_2H_2



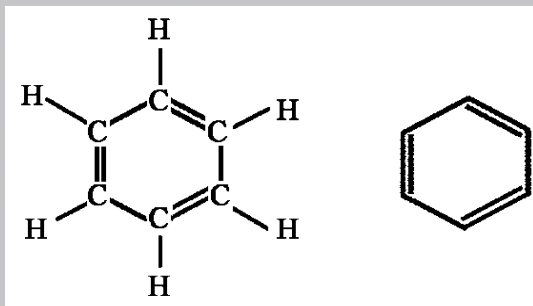
Lewis structure



LINEAR
angle 180°
molecular structure



AROMATIC HYDROCARBONS

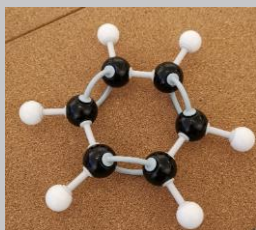
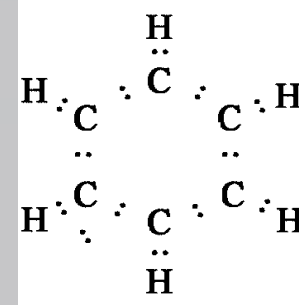


benzene
 C_6H_6

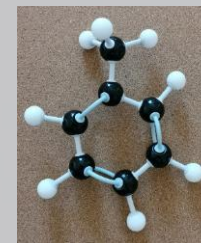
carbon - $1s^2 2s^2 2p^2$ (4 electrons in the valence shell)

3 of these electrons form σ bonds to 2 carbons and 1 hydrogen.

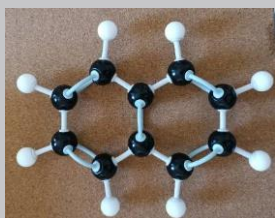
The “spare electrons” are in orbitals above and below the plane of the molecule forming π bonds.



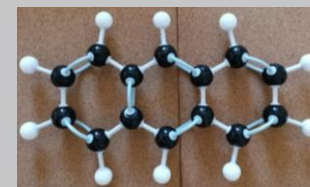
benzene, C_6H_6 ; m.p. $5.6^\circ C$ b.p. $80.1^\circ C$



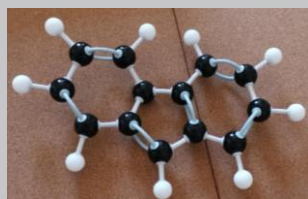
toluene, C_7H_8 ; m.p. $-95.0^\circ C$ b.p. $110.6^\circ C$



naphthalene, $C_{10}H_8$; m.p. $80.3^\circ C$ b.p. $218^\circ C$



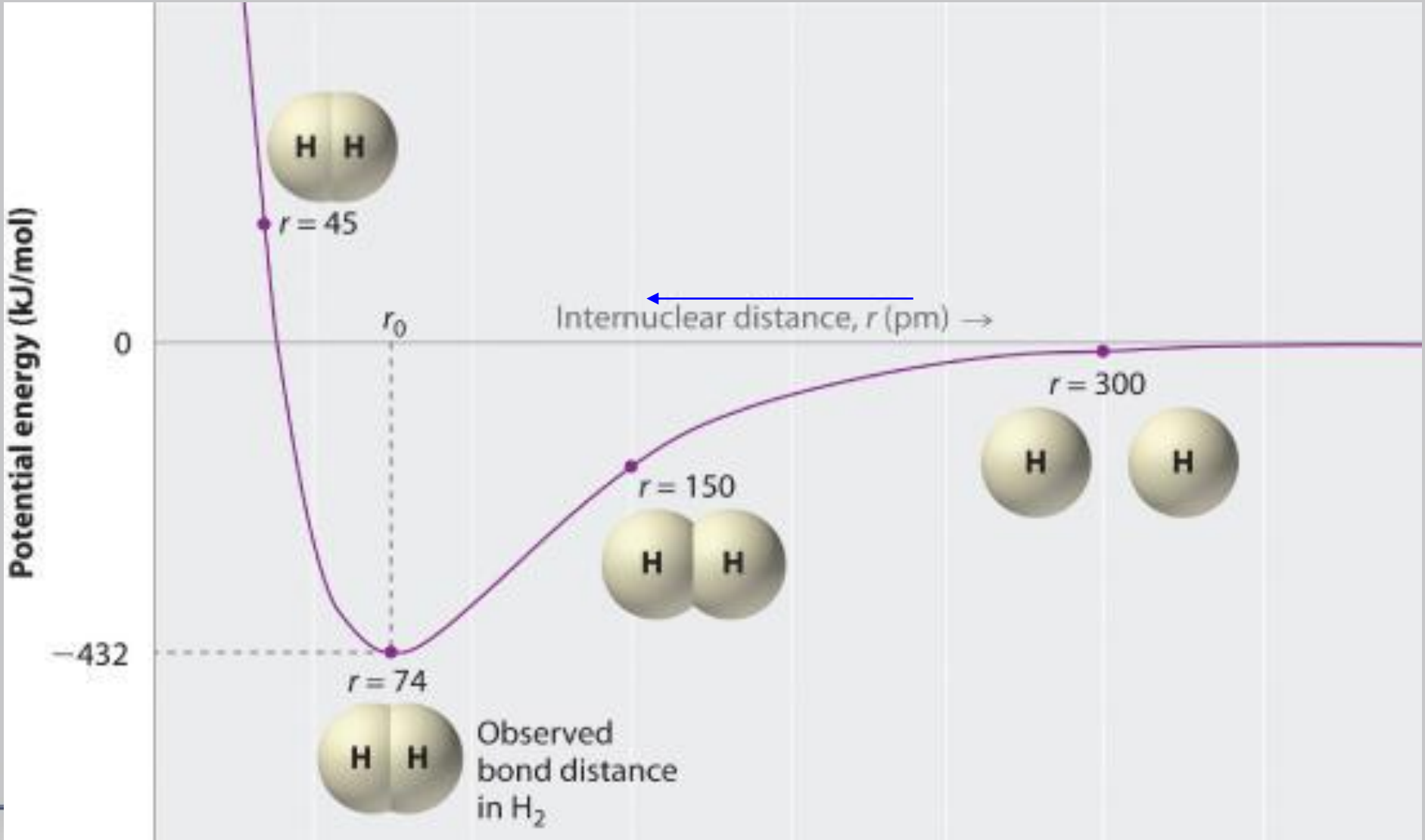
anthracene, $C_{14}H_{10}$; m.p. $216^\circ C$ b.p. $354^\circ C$



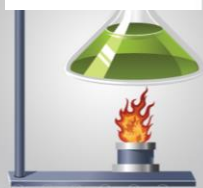
fluorene, $C_{13}H_{10}$; m.p. $114^\circ C$ b.p. $295^\circ C$



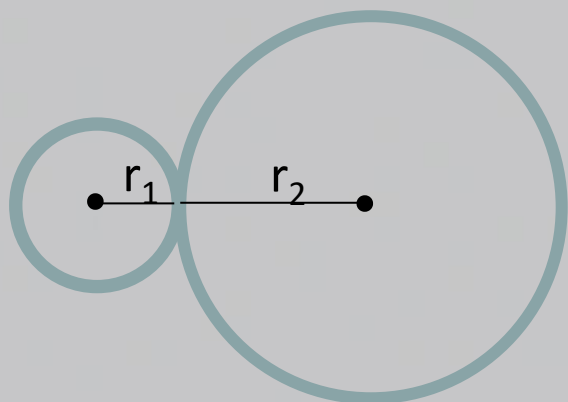
COVALENT BONDS



bond formation – energy is released !!!



BOND LENGTHS (distance)



BOND LENGTH - sum of the atomic radii $r_1 + r_2$

X-Y	BL (pm)	X-Y	BL (pm)	X-Y	BL (pm)
C-H	109	N-H	101	O-H	96
C-Cl	177	N-Cl	191	Cl-H	127
C-Br	194	N-Br	214	Br-H	141
C-I	214	N-I	222	I-H	161

The bigger the atoms, the longer the bond

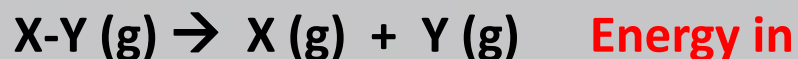
X-Y	BL (pm)	X-Y	BL (pm)	X-Y	BL (pm)
C-C	154	N-N	145	O-O	141
C=C	134	N=N	122	O=O	121
C≡C	120	N≡N	110	H-H	74

**Single bond longer than double;
double longer than triple**



BOND ENERGY (strength)

BOND STRENGTH - energy that has to be put in to take two bonded atoms apart
OR energy that is given out when two atoms bond together



X-Y	$D_{\text{X-Y}}$ (kJ/mol)	X-Y	$D_{\text{X-Y}}$ (kJ/mol)	X-Y	$D_{\text{X-Y}}$ (kJ/mol)
C-H	413	N-H	391	O-H	366
C-Cl	330	N-Cl	200	Cl-H	432
C-Br	228	N-Br	243	Br-H	366
C-I	214	N-I	159	I-H	298

In general, the longer the bond, the weaker the bond

X-Y	BE (kJ/mol)	X-Y	BE (kJ/mol)	X-Y	BE (kJ/mol)
C-C	348	N-N	170	O-O	145
C=C	614	N=N	418	O=O	498
C≡C	839	N≡N	945		

Single weaker than double;
double weaker than triple



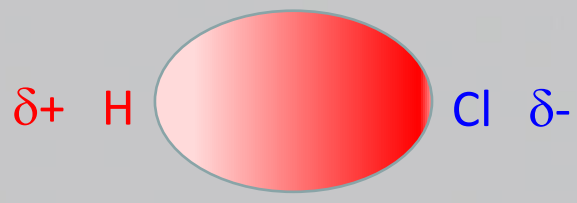
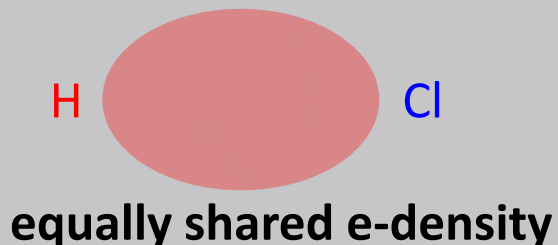
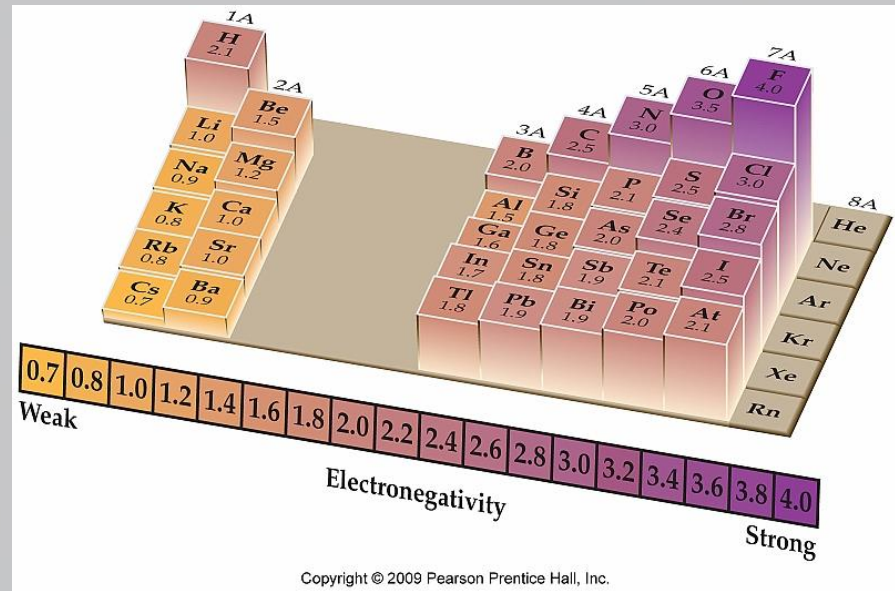
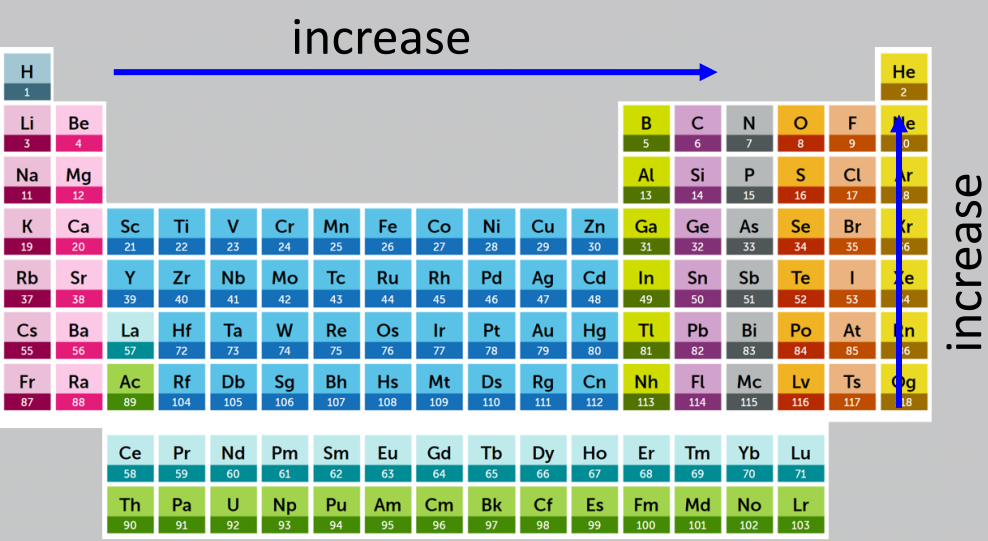
ELECTRONEGATIVITY

The measure of the extent to which an atom desires electrons in a shared bond X--Y

If Y is more electronegative than X, Y pulls electrons more.

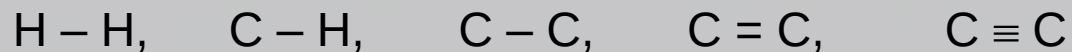
Unequal sharing electrons Partial ‘ionic character’

Less electronegative atom gets a little more positive



POLAR BOND

QUESTION 1: Predict which of the following bonds would be the longest?



QUESTION 2: Which atom is more electronegative, oxygen or selenium?

QUESTION 3: In which of the following bonds would the oxygen atom would be the slightly positive one?

