The periodic table

MAIN-GROUP ELEMENTS (groups 1A–2A) · TRANSITION ELEMENTS (groups 3B–2B) · MAIN-GROUP ELEMENTS (groups 3A–8A)

Period	1A (1)	2A (2)	3B (3)	4B (4)	5B (5)	6B (6)	7B (7)	8B (8)	8B (9)	8B (10)	1B (11)	2B (12)	3A (13)	4A (14)	5A (15)	6A (16)	7A (17)	8A (18)
1	1 **H** 1.008																	2 **He** 4.003
2	3 **Li** 6.941	4 **Be** 9.012											5 **B** 10.81	6 **C** 12.01	7 **N** 14.01	8 **O** 16.00	9 **F** 19.00	10 **Ne** 20.18
3	11 **Na** 22.99	12 **Mg** 24.31											13 **Al** 26.98	14 **Si** 28.09	15 **P** 30.97	16 **S** 32.07	17 **Cl** 35.45	18 **Ar** 39.95
4	19 **K** 39.10	20 **Ca** 40.08	21 **Sc** 44.96	22 **Ti** 47.90	23 **V** 50.94	24 **Cr** 52.00	25 **Mn** 54.94	26 **Fe** 55.85	27 **Co** 58.93	28 **Ni** 58.70	29 **Cu** 63.55	30 **Zn** 65.39	31 **Ga** 69.72	32 **Ge** 72.59	33 **As** 74.92	34 **Se** 78.96	35 **Br** 79.90	36 **Kr** 83.80
5	37 **Rb** 85.47	38 **Sr** 87.62	39 **Y** 88.91	40 **Zr** 91.22	41 **Nb** 92.91	42 **Mo** 95.94	43 **Tc** (98)	44 **Ru** 101.1	45 **Rh** 102.9	46 **Pd** 106.4	47 **Ag** 107.9	48 **Cd** 112.4	49 **In** 114.8	50 **Sn** 118.7	51 **Sb** 121.8	52 **Te** 127.6	53 **I** 126.9	54 **Xe** 131.3
6	55 **Cs** 132.9	56 **Ba** 137.3	57 **La** 138.9 / 72 **Hf** 178.5	73 **Ta** 180.9	74 **W** 183.9	75 **Re** 186.2	76 **Os** 190.2	77 **Ir** 192.2	78 **Pt** 195.1	79 **Au** 197.0	80 **Hg** 200.6	81 **Tl** 204.4	82 **Pb** 207.2	83 **Bi** 209.0	84 **Po** (209)	85 **At** (210)	86 **Rn** (222)	
7	87 **Fr** (223)	88 **Ra** (226)	89 **Ac** (227) / 104 **Unq** (261)	105 **Unp** (262)	106 **Unh** (263)	107 **Uns** (262)	108 **Uno** (265)	109 **Une** (267)										

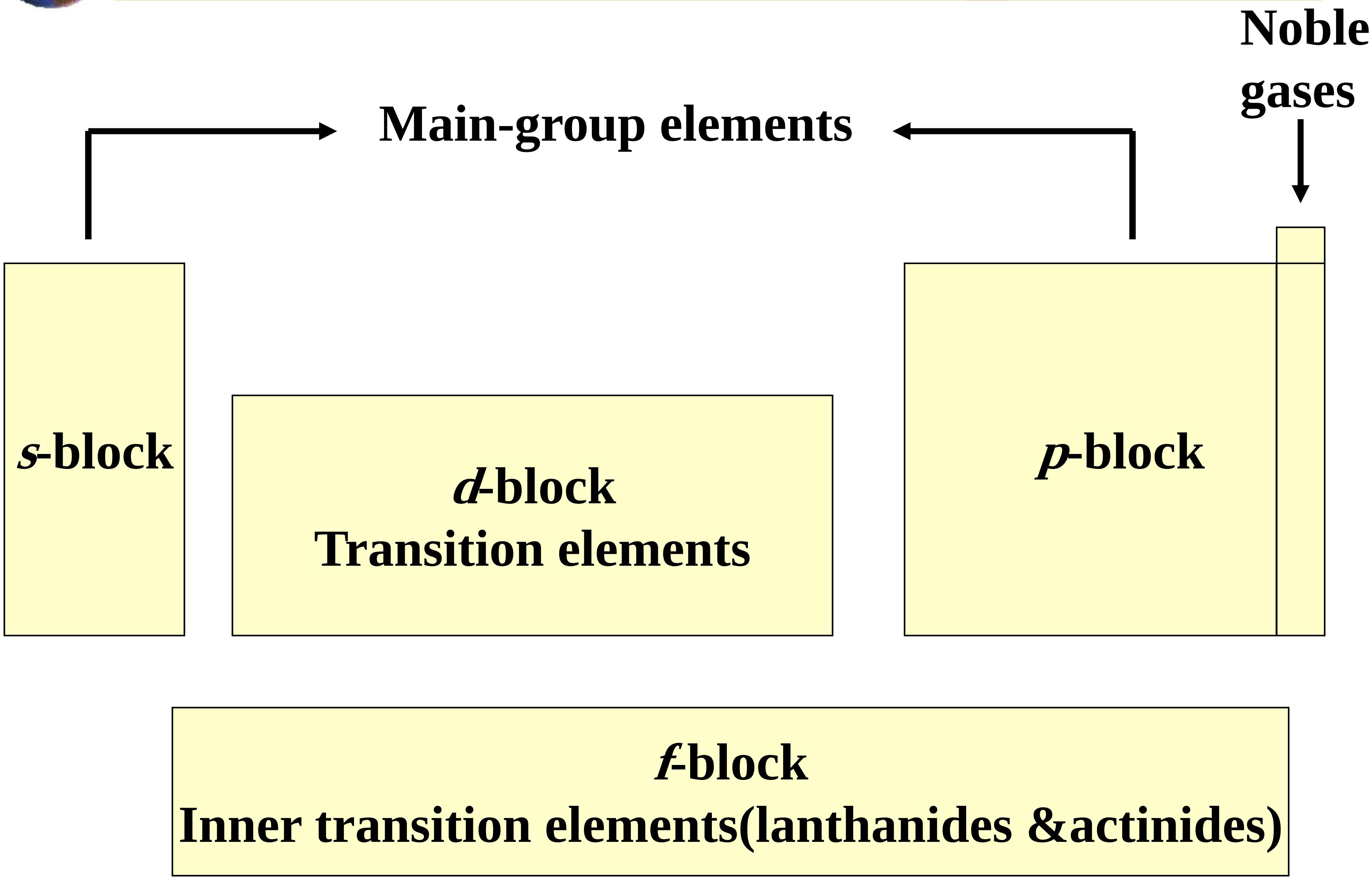
Period

Non-metals

INNER TRANSITION ELEMENTS

6	Lanthanides	58 **Ce** 140.1	59 **Pr** 140.9	60 **Nd** 144.2	61 **Pm** (145)	62 **Sm** 150.4	63 **Eu** 152.0	64 **Gd** 157.3	65 **Tb** 158.9	66 **Dy** 162.5	67 **Ho** 164.9	68 **Er** 167.3	69 **Tm** 168.9	70 **Yb** 173.0	71 **Lu** 175.0
7	Actinides	90 **Th** 232.0	91 **Pa** (231)	92 **U** 238.0	93 **Np** (244)	94 **Pu** (242)	95 **Am** (243)	96 **Cm** (247)	97 **Bk** (247)	98 **Cf** (251)	99 **Es** (252)	100 **Fm** (257)	101 **Md** (258)	102 **No** (259)	103 **Lr** (260)

Classification of elements according to electronic configuration



Using the periodic table to write the electronic configuration

Fe $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^6$

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

Fe: $[Ar]\ 4s^2\ 3d^6$

$_{18}Ar$: third period, next subshell to be filled 4s

Example:

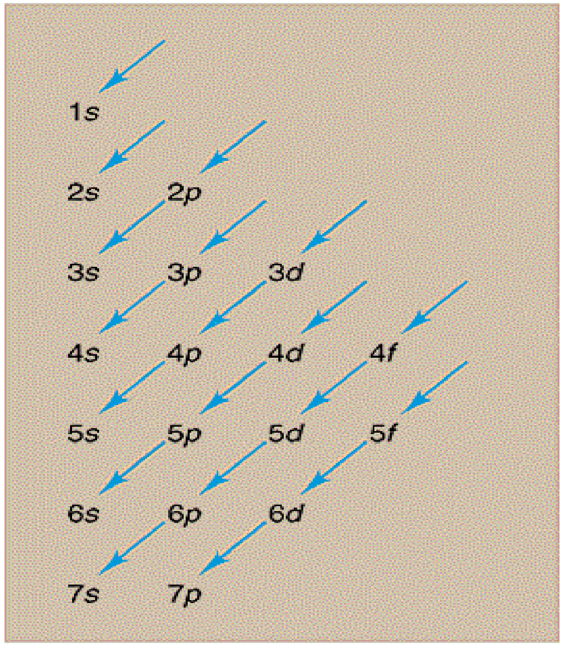
write the electronic configuration of $_{60}Nd$

$_{54}Xe$, is the noble gas before Nd

6 electrons needed to reach Nd

Xe: 5th period

Nd: $[Xe]6s^2 4f^4$



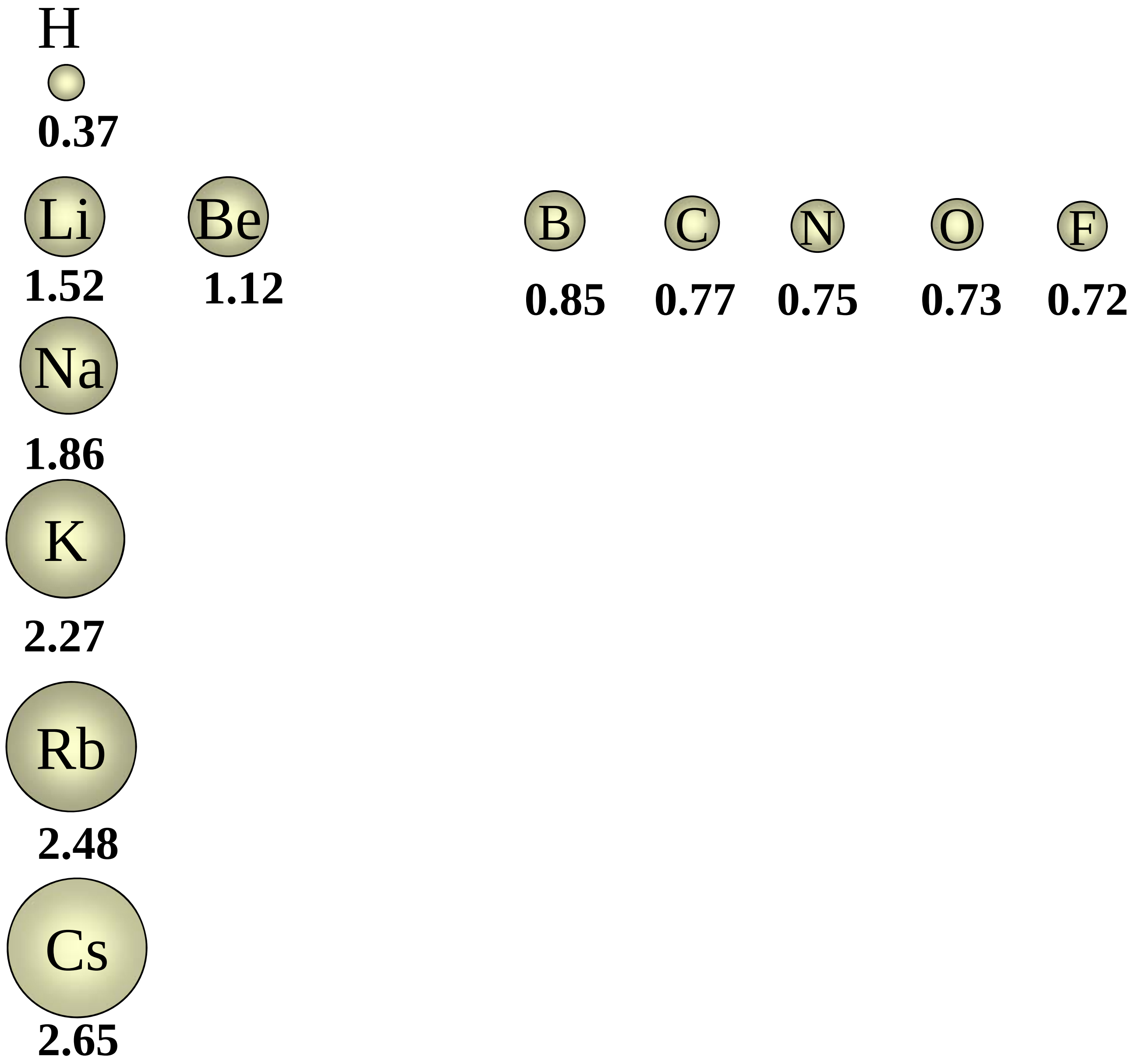
Periodic properties of atoms

How an atom reacts depends on many factors:
nuclear charge, electronic configuration, volume,….

Atomic sizes:
In a group atomic radius increases from top to bottom
In a period atomic radius decreases from left to right

Valence orbital: last shell filled (highest n)

Atomic radii



Unit:Å

Ionization energies

Ionization energy: minimum energy required to remove an electron from a gaseous atom in its ground state.

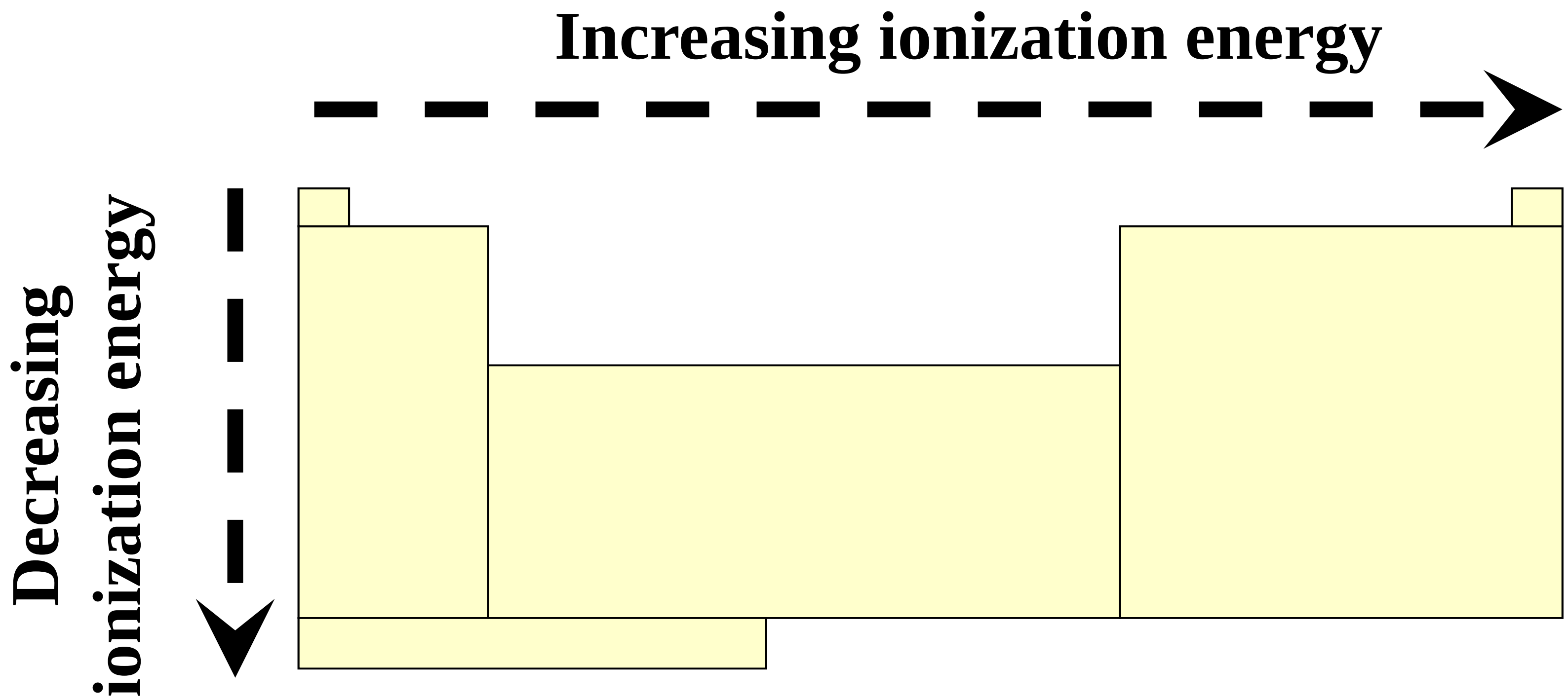
$$A(g) \rightarrow A^{+}(g) + e^{-}$$

Usually: a unit of eV is used for one electron and kJ/mol for one mole of atoms

In general, ionization energy increases across a period from left to right

In general, ionization energy decreases within a group of main groups elements from top to bottom

Ionization energy



Each of The following elements has an ionization energy higher than the ionization energy of the element that follows it.

- **The noble gases(He, Ne, Ar, Kr, Xe, and Rn)**
  Electronic configuration: ……..$ns^2\ np^6$

- **The elements Be, Mg, Zn ,Cd and Hg, each of which has a filled s subshell in the outermost shell**
  Electronic configuration: ……..ns^2

- **The elements N, P, and As, each of which has a half-filled p subshell in the outer most shell**
  Electronic configuration: ……..$ns^2\ np^3$

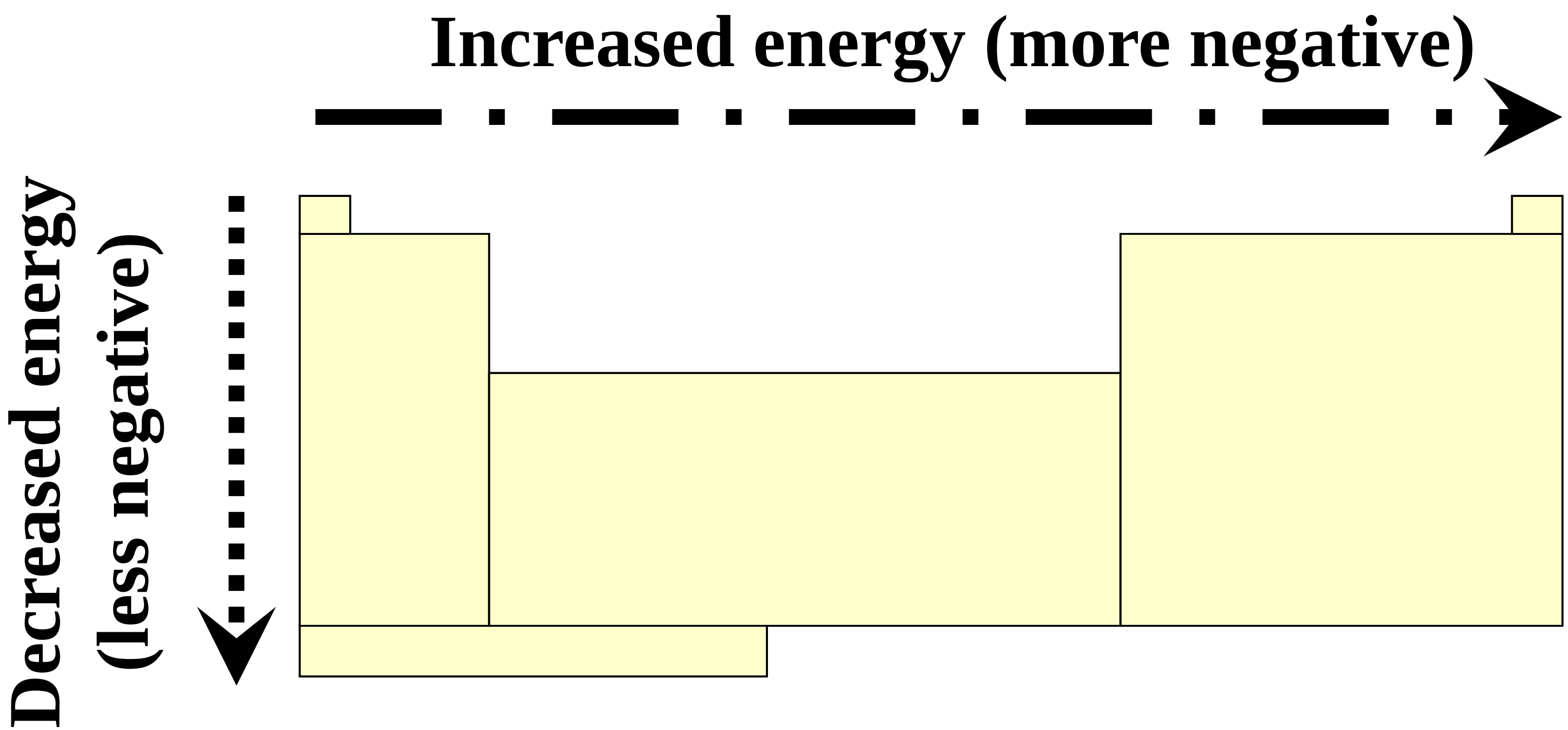
for positive ions more energy is needed for ionization

Third I.E. > Second I.E. > First I.E.

Electron affinities

Electron affinity: the energy change when an electron is added to a gaseous atom in its ground state

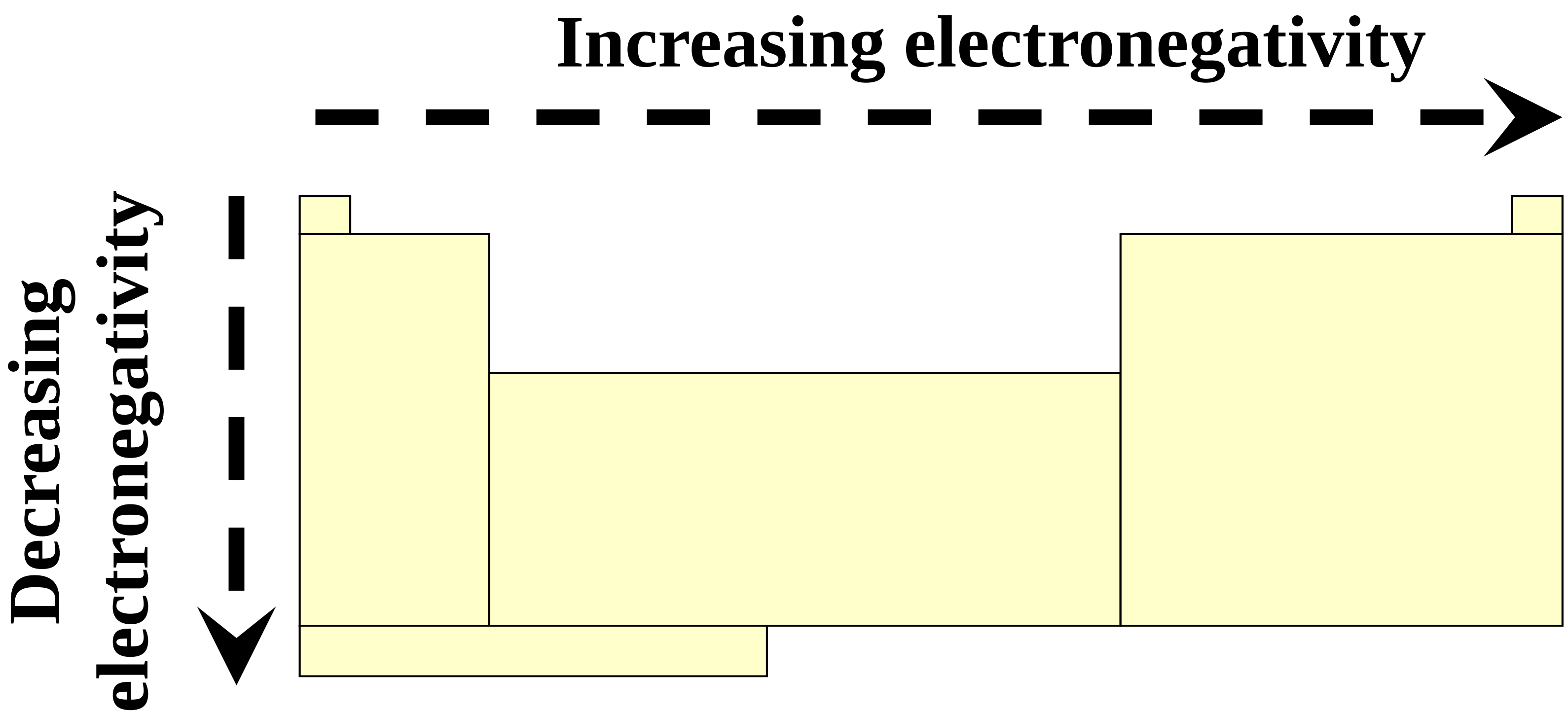
$$e^- + A(g) \rightarrow A^-(g)$$



Exceptions to this generality should be noted (same as I.E.)

Electronegativity

Electronegativity: a measure of the relative tendency of an atom to attract electrons to itself.



Examples

Arrange the following elements in order of increasing electro negativity: B, Na, F, O

Na < B < O < F

Arrange the following elements in order of decreasing ionization energy: P, N, O, F

F > N > O > P

Determine the largest atom among the following elements: Sn, Ba, Al, Ga

Ba has the largest radius

Chemical bonds

The ionic bond: metal + nonmetal: electrons are transferred from atoms of the metal to the atoms of the nonmetal

Cations: the atoms that lose electrons
Anions: the atoms that gain electrons

These ions attract one another to form a crystal

Consider the reaction of the a sodium atom with a chlorine atom

$$\mathrm{Na}\ (1s^2\ 2s^2\ 2p^6\ 3s^1) \rightarrow \mathrm{Na}^+\ (1s^2\ 2s^2\ 2p^6) + e^-$$

$$e^- + \mathrm{Cl}\ (1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5) \rightarrow \mathrm{Cl}^-\ (1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6)$$

$$Na + Cl \rightarrow Na^{+} + Cl^{-}$$

The sodium ion has an electronic configuration identical to that of neon

The chloride ion has an electronic configuration identical to that of argon

The covalent bond

When atoms of nonmetals interact,
molecules formed are held together by covalent bonds

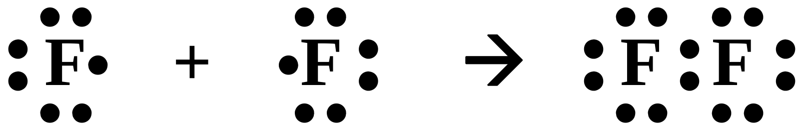
Consider the molecule of hydrogen H_2, both atoms of hydrogen are similar in their attraction of electrons electron transfer does not occur instead electrons are shared

H : H

Each hydrogen atom has a configuration similar to He

H — H

Elements of group 7A form molecules held together by covalent bonds

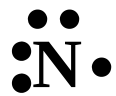


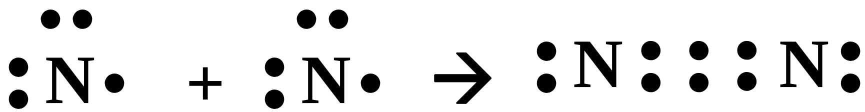
F_2

Cl_2, Br_2, and I_2 follow the same pattern

More than one covalent bond may form between two atoms

Nitrogen atom has five valence electrons





N_2

The electron-dot formulas are called *Lewis Structures*

Lewis theory:
a noble gas configuration is attained in covalent bonded atoms.

For nonmetals: number of valence electrons = group number

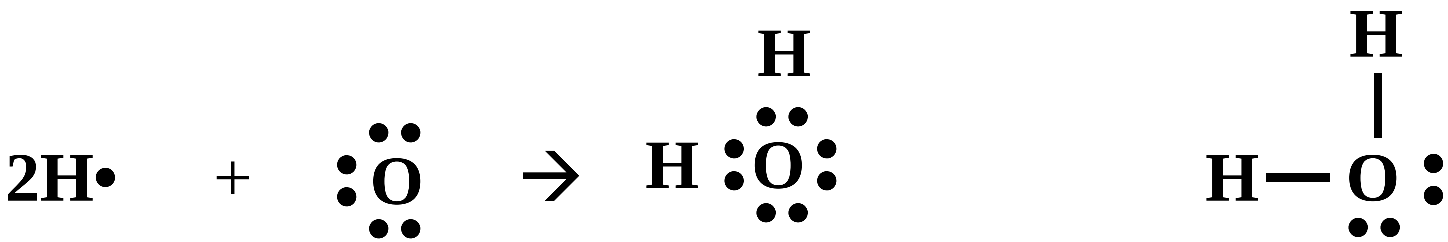
Prediction: to attain a stable octet

VII A elements, such as Cl, would form one covalent bond

VI A elements, such as O, would form two covalent bonds

V A elements, such as N, would form three covalent bonds

Hydrogen chloride



water

Systematic procedure for drawing Lewis structures

Example: ClO_3^-

(a) Determine the total number of valence electrons in the molecule

main groups elements: valence electrons = group number

If the ion has a negative charges:
add the value of the charge to total

If the ion has a positive charges:
subtract the value of the charge from total

(a) **Total valence electrons = 7+ 3×6 + 1 = 26**

(b) Determine the number electrons that would be required to give 2 electrons to each hydrogen atom and 8 electrons to each of the other atoms individually.

For ClO_3^-

(b) total number of electrons required for individual atoms

4 × 8 = 32 electrons

(c) Determine the number of electrons that must be shared in the final structure

This would be the number resulted from: step (b) minus step *(a)*

For ClO_3^-

(c) Number of shared electrons = 32 – 26 = 6

(d) determine the number of covalent bonds in the molecule

This would be the number resulted from: step (c)/ 2

For ClO_3^-

(d) Number of covalent bonds = 6/2 =3 bonds

(e) Determine the number of unshared electrons in the molecule

This would be the number resulted from:

valence electrons {step (a)} – shared electrons {step (c)}

For ClO_3^-

(e) Number of unshared electrons = 26 – 6 = 20 electrons

(f) Draw the structural formula

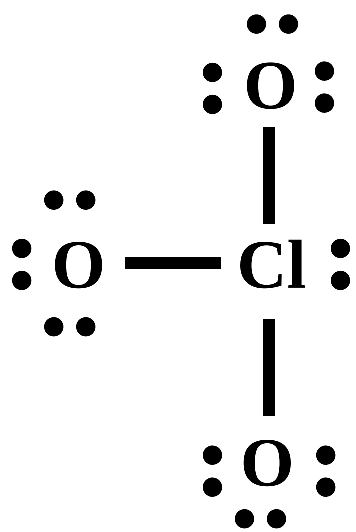
Write down the chemical symbols for the atoms

Draw a covalent bond between each atom

Draw multiple bonds as needed (hydrogen: one covalent bond only)

Assign unshared electrons to the atoms bringing a total of 8 around each atom (H: 2)

For ClO_3^-



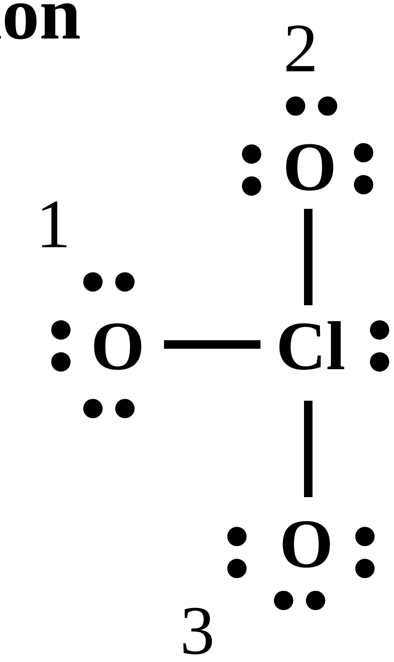
(g) Determine the formal charge on each atom

Formal charge=group number – [#bonds + #unshared electrons]

For neutral molecules: ∑ formal charges = 0

For ions: ∑ formal charges = charge of the ion

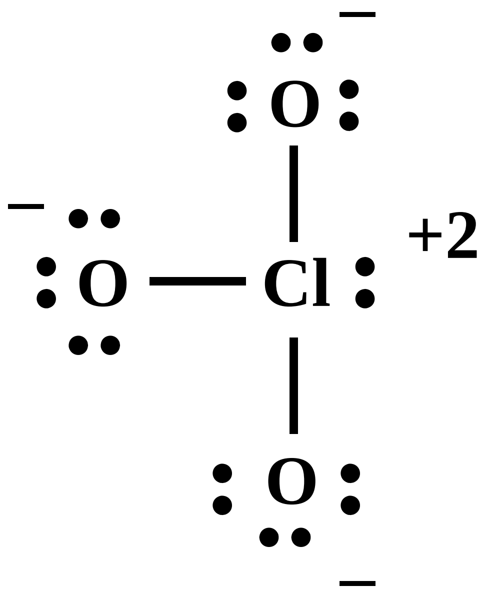
For ClO_3^-



F.C. (for each oxygen) = 6 – [1 + 6] = –1

F.C. (for Cl) = 7 – [3 + 2] = +2

Lewis structure for ClO_3^-



Example

Diagram the Lewis structure of SO_2(S: the central atom)

(a) Total valence electrons = 6 × 1+ 6 × 2 = 18

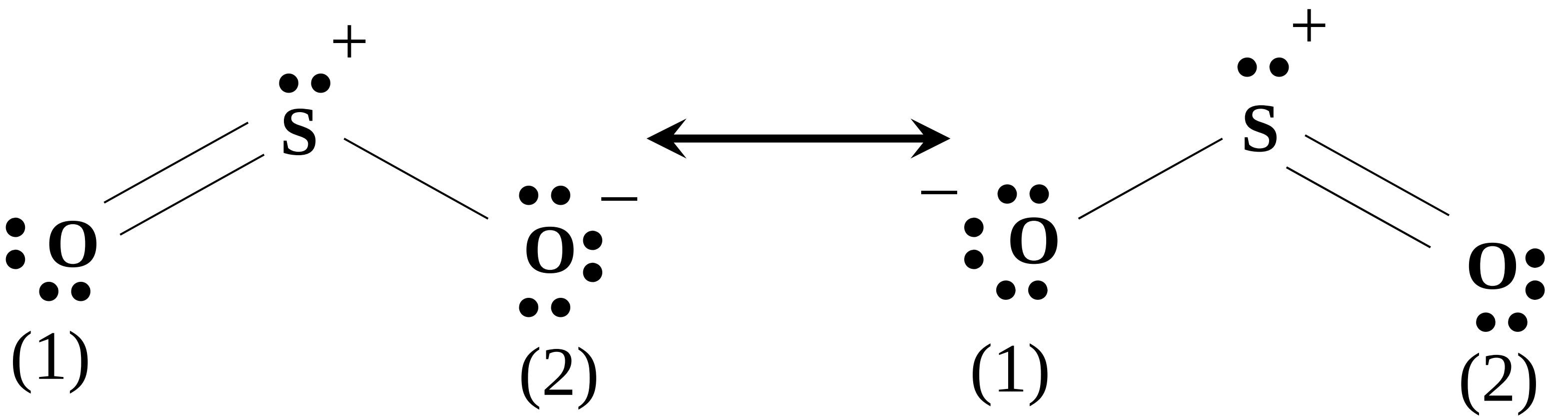
(b) # electrons required for individual atoms 3 × 8 = 24 e^-s

(c) Number of shared electrons = 24 – 18 = 6

(d) Number of covalent bonds = 6/2 =3 bonds

(e) Number of unshared electrons = 18 – 6 = 12 electrons

(f) Draw the structural formula



F.C. (S) = 6 – [3 + 2] = +1

F.C. $(O)_1$ = 6 – [2+ 4] = 0

F.C. $(O)_2$ = 6 – [1+ 6] = – 1

F.C. (S) = 6 – [3 + 2] = +1

F.C. $(O)_1$ = 6 – [2+ 4] = – 1

F.C. $(O)_2$ = 6 – [1+ 6] = 0

The two structures are said to be resonance forms of SO_2

Example

Diagram the Lewis structure of CO_3^{-2}
(C: the central atom)

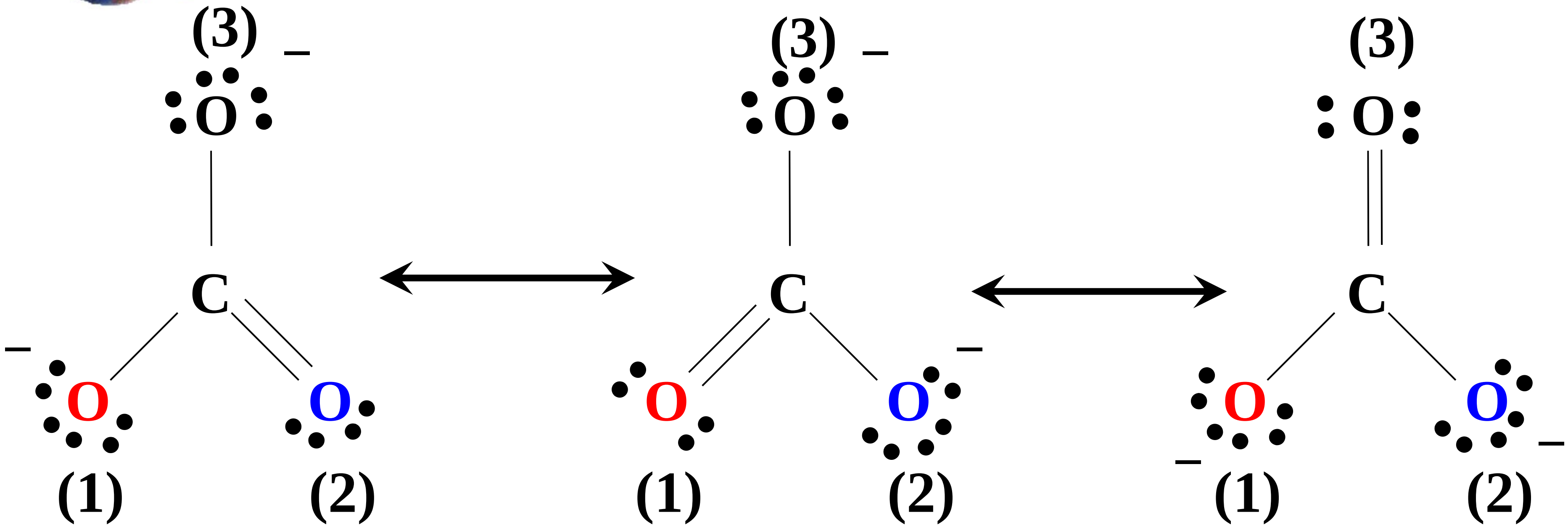
(a) Total valence electrons = 4×1 + 6×3 + 2 = 24

(b) # electrons required for individual atoms 4 × 8 = 32

(c) Number of shared electrons = 32 – 24 = 8

(d) Number of covalent bonds = 8/2 = 4 bonds

(e) Number of unshared electrons = 24 – 8 = 16 electrons



$FC(C) = 4 – [4] = 0$

$FC(O)_1 = 6–[1+6]= –1$

$FC(O)_2 = 6–[2+4] = 0$

$FC(O)_3 = 6–[1+6]= –1$

$FC(C) = 4 – [4] = 0$

$FC(O)_1 = 6–[2+4]=0$

$FC(O)_2 = 6–[1+6] = –1$

$FC(O)_3 = 6–[1+6]= –1$

$FC(C) = 4 – [4] = 0$

$FC(O)_1 = 6–[1+6]= –1$

$FC(O)_2 = 6–[2+4] = 0$

$FC(O)_3 = 6–[1+6]= –1$

Example

Diagram the Lewis structure of N_2O (the atoms are arranged NNO)

(a) Total valence electrons = 5×2 + 6×1 = 16

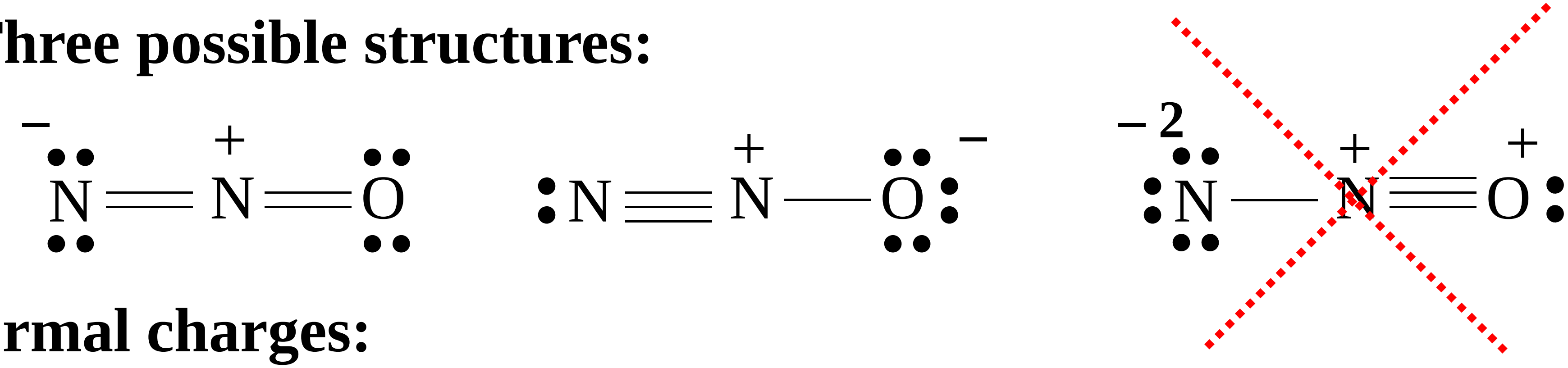
(b) # electrons required for individual atoms 3 × 8 = 24

(c) Number of shared electrons = 24 – 16 = 8

(d) Number of covalent bonds = 8/2 = 4 bonds

(e) Number of unshared electrons = 16 – 8 = 8 electrons

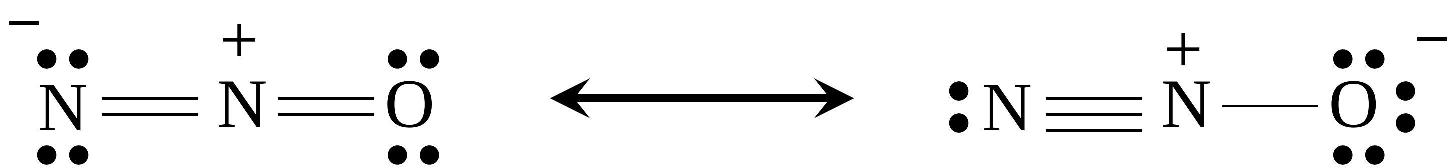
Three possible structures:



Formal charges:

$(N)_1 = 5 - [2+4] = -1$	$(N)_1 = 5 - [3+2] = 0$	$(N)_1 = 5 - [1+6] = -2$
$(N)_2 = 5 - [4+0] = +1$	$(N)_2 = 5 - [4+0] = +1$	$(N)_2 = 5 - [4+0] = +1$
$(O) = 6 - [2+4] = 0$	$(O) = 6 - [1+6] = -1$	$(O) = 6 - [3+2] = +1$

Resonance forms:



Example:

diagram the resonance forms for the NPNH molecule

(a) Total valence electrons = 5 + 5 + 5 + 1 = 16

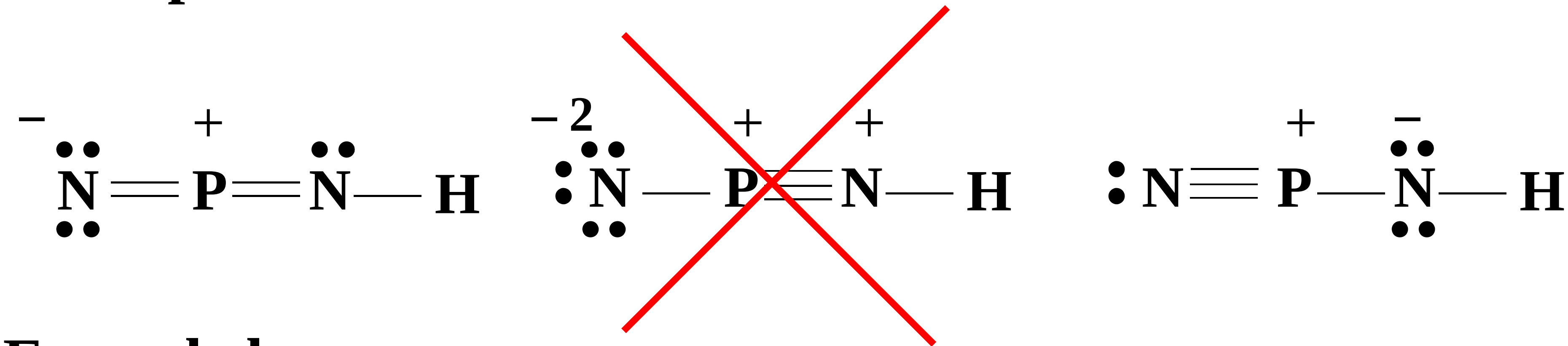
(b) # electrons required for individual atoms 3 × 8 + 2= 26

(c) Number of shared electrons = 26 – 16 = 10

(d) Number of covalent bonds = 10/2 = 5 bonds

(e) Number of unshared electrons = 16 – 10 = 6 electrons

Three possible structures



Formal charges:

$(N)_1 = 5 - [2+4] = -1$

$(P) = 5 - [4+0] = +1$

$(N)_2 = 5 - [3+2] = 0$

$(H) = 1 - [1] = 0$

$(N)_1 = 5 - [1+6] = -2$

$(P) = 5 - [4+0] = +1$

$(N)_2 = 5 - [4+0] = +1$

$(H) = 1 - [1] = 0$

$(N)_1 = 5 - [3+2] = 0$

$(P) = 5 - [4+0] = +1$

$(N)_2 = 5 - [2+4] = -1$

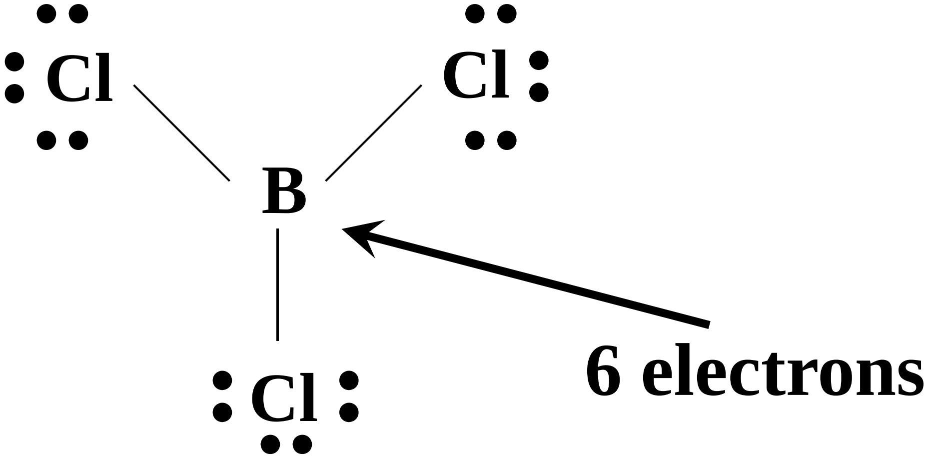
$(H) = 1 - [1] = 0$

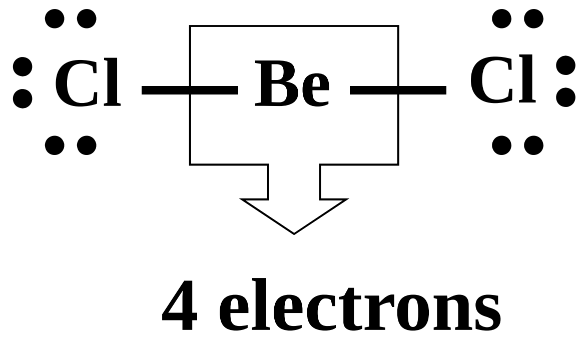
Resonance forms:

Exceptions to the octet rule

Some stable molecules exist that does not have noble gas configurations

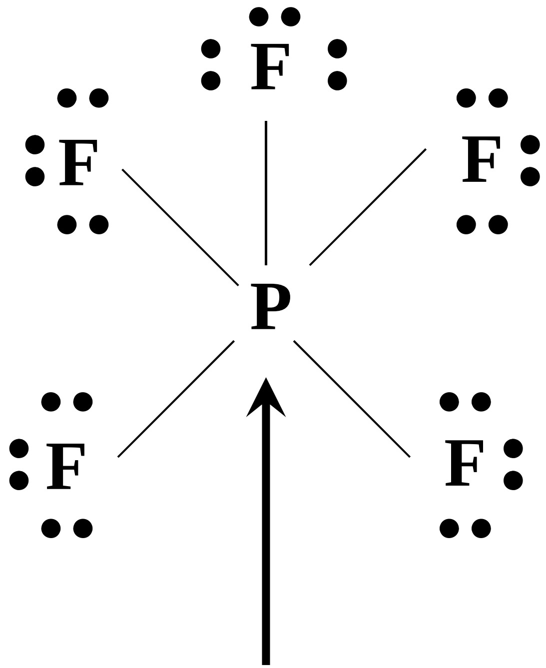
➢**Molecules contain atoms with less than 8 valence shell electrons**





$BeCl_2$

➢**Molecules contain atoms with more than 8 valence shell electrons**



10 electrons

SF_6

How many valence shell electrons are there in S?

- **Molecules contain atoms with odd number of valence electrons**

NO	**NO_2**
Nitrous oxide	**Nitrogen dioxide**