LEWIS STRUCTURES

LEWIS THEORY OF BONDING

-Atoms & ions are stable if they have a noble-gas like electron structure (stable octet)

-Electrons are most stable when paired

-Atoms form chemical bonds to achieve a stable octet of electrons

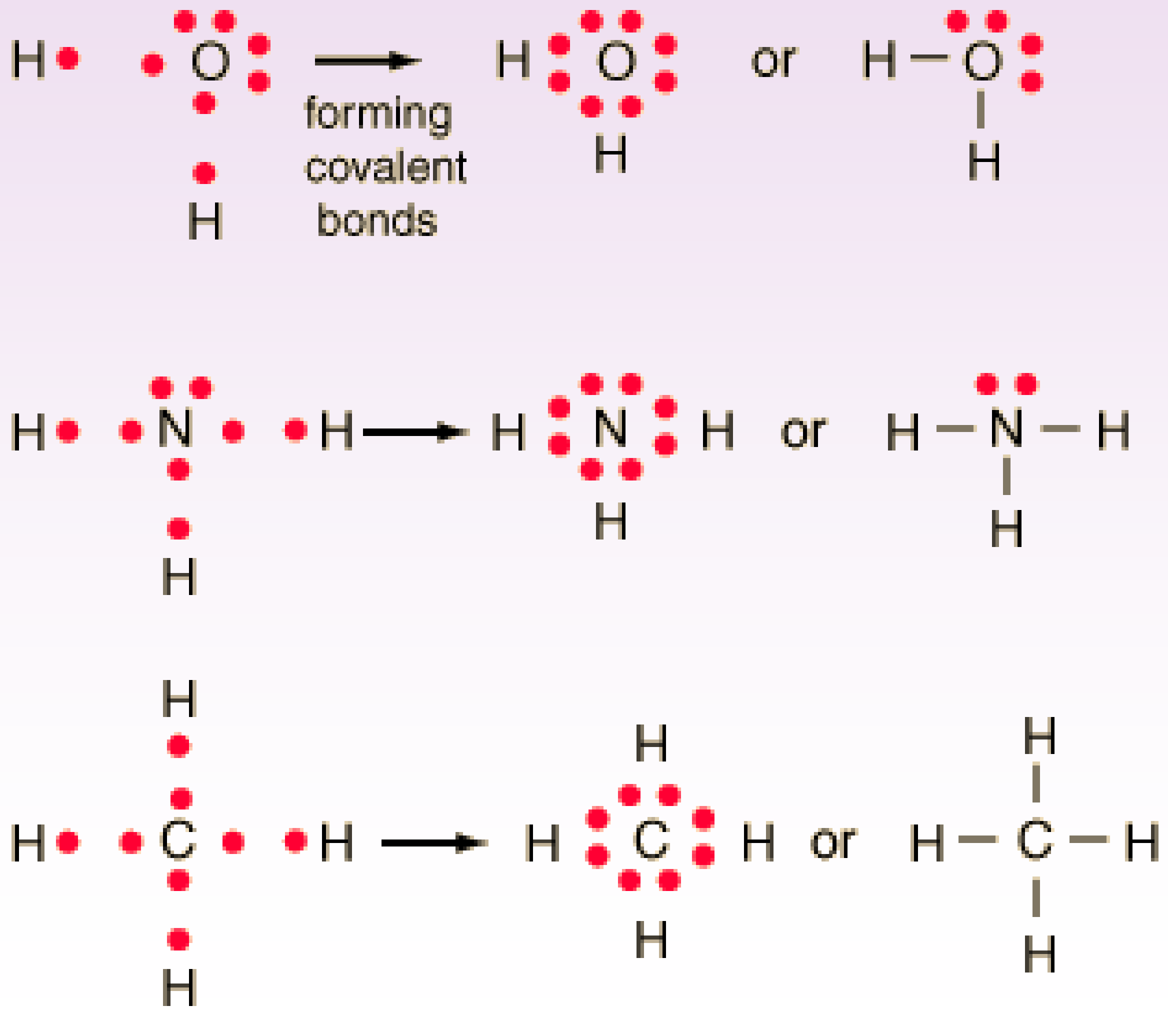
-A stable octet may be achieved by:

-An exchange of electrons between metal and non-metal atoms (ionic bonding)

-A sharing of electrons between 2 non-metal atoms (covalent bonding)

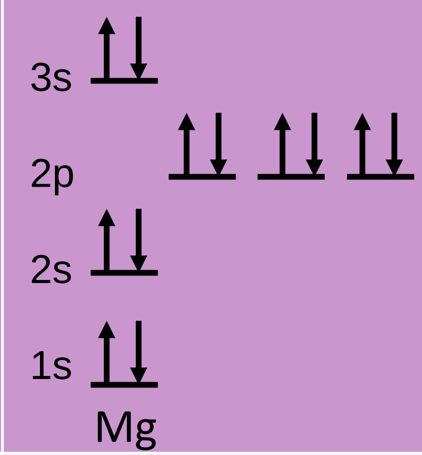
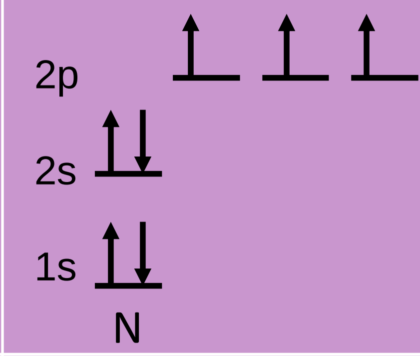
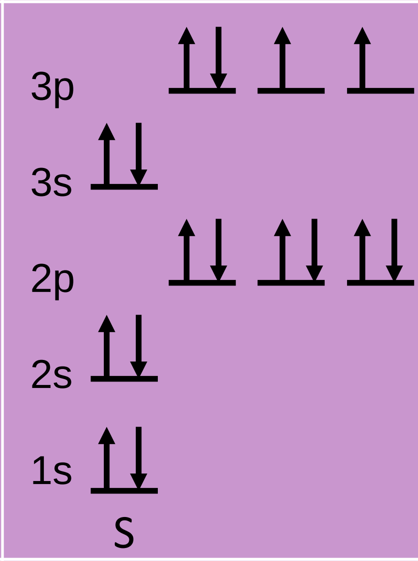
LEWIS STRUCTURES

A Lewis structure communicates the arrangement of electrons and bonds in a chemical substance



LEWIS STRUCTURES

Lewis Dot Diagrams correspond to Hund's rule governing the placement of electrons into orbitals

Element	Mg	N	S
Valence	2+	3-	2-
Lewis Symbol	$\dot{\mathrm{Mg}}\cdot$	$\underset{\cdot}{\cdot\ddot{\mathrm{N}}\cdot}$	$\underset{\cdot}{\cdot\ddot{\mathrm{S}}:}$
Energy-Level Diagram			
Electron Configuration	Mg: $1s^2\ 2s^2\ 2p^6\ 3s^2$	N: $1s^2\ 2s^2\ 2p^3$	S: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$

DRAWING LEWIS STRUCTURES

Draw: H_2O

H O H

1. Decide atom arrangement
2. Count ALL valence electrons (add electrons and subtract electrons based on the molecule's charge)
3. Place 2 electrons in each bond
4. Complete the octets of the atoms attached to the centre atom by adding electrons in pairs
5. Place remaining electrons on the central atom in pairs
6. If the central atom does not have an octet, form double or triple bonds
7. Use formal charges to determine if more double or triple bonds can be added to the central atom (if in 3rd energy level or higher):

Formal charge = Valence e^- - lone pair e^- - # of bonds

DRAWING LEWIS STRUCTURES

Draw: H_2O

$$\mathbf{H} \quad \mathbf{O} \quad \mathbf{H}$$

1 + 6 + 1 = 8 electrons

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DRAWING LEWIS STRUCTURES

Draw: H_2O

H:O:H

4 electrons remaining

1+6+1~~=8 electrons~~

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DRAWING LEWIS STRUCTURES

Draw: H_2O

$$H:O:H$$

4 electrons remaining

The octets on hydrogen are already full.

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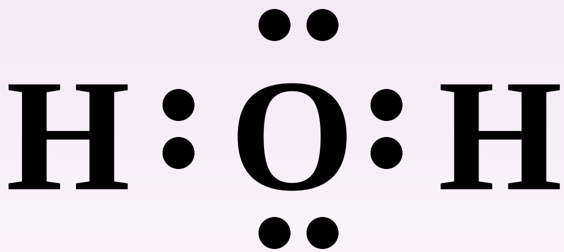
DRAWING LEWIS STRUCTURES

Draw: H_2O



0 electrons

~~4 electrons~~ remaining

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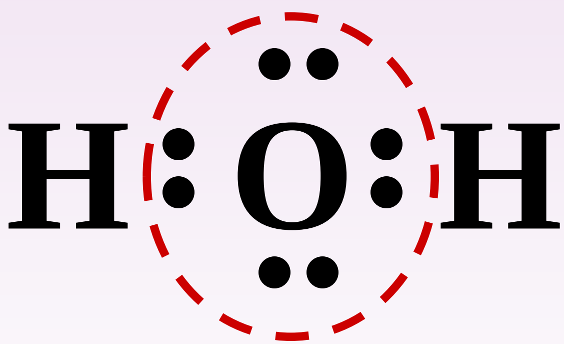
DRAWING LEWIS STRUCTURES

Draw: H_2O



0 electrons remaining

Oxygen has a full octet, so there is no need to form double or triple bonds in this case

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DRAWING LEWIS STRUCTURES

Draw: SO_2

O S O

1. Decide atom arrangement
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DRAWING LEWIS STRUCTURES

Draw: SO_2

O S O

6 + 6 + 6 = 18 electrons

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Formal charge = Valence e^- - lone pair e^- - # of bonds

DRAWING LEWIS STRUCTURES

Draw: SO_2

$$O : S : O$$

14 electrons remaining

6+6+6= ~~18 electrons~~

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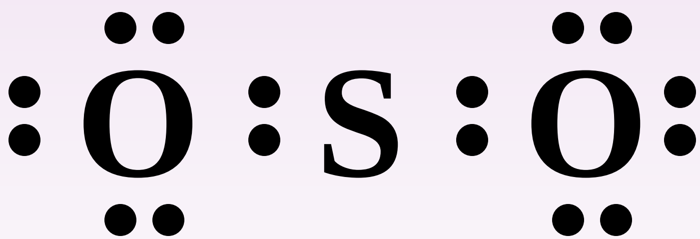
DRAWING LEWIS STRUCTURES

Draw: SO_2



2 electrons
~~14 electrons~~ remaining

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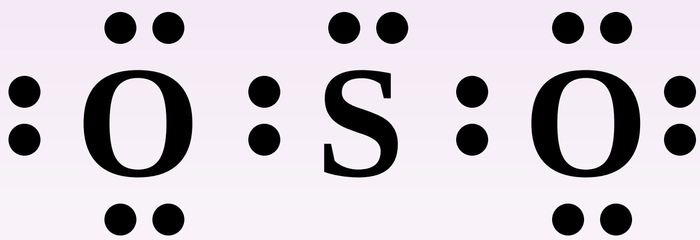
DRAWING LEWIS STRUCTURES

Draw: SO_2



~~2 electrons~~

0 electrons

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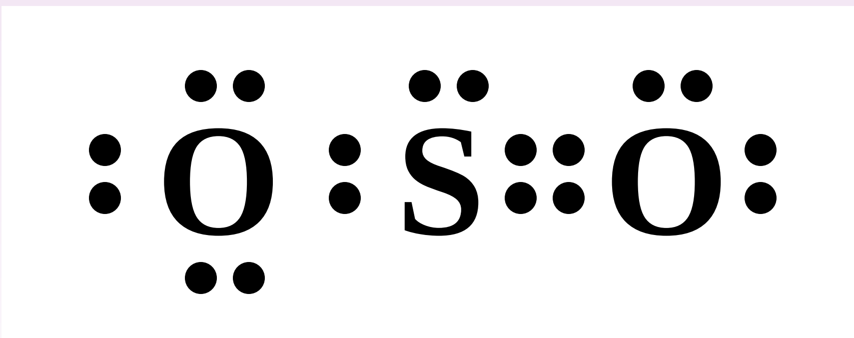
DRAWING LEWIS STRUCTURES

Draw: SO_2



0 electrons

Now all atoms have a full octet

1. Decide atom arrangement
2. Count ALL valence electrons (add electrons and subtract electrons based on the molecule's charge)
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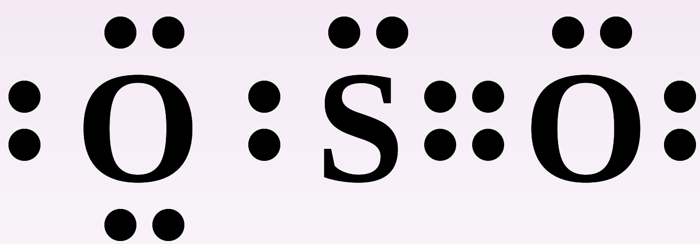
DRAWING LEWIS STRUCTURES

Draw: SO_2

6 - 2 - 3 = +1



6 - 6 - 1 = -1

6 - 4 - 2 = 0

There should be a formal charge of zero for each atom in a neutral molecule.

This means that this diagram needs to be modified.

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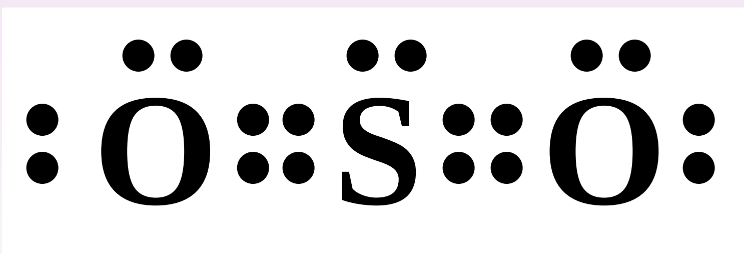
7. Use formal charges to determine if more double or triple bonds can be added to the central atom (if in 3rd energy level or higher):

Formal charge = Valence e^- - lone pair e^- - # of bonds

DRAWING LEWIS STRUCTURES

Draw: SO_2

6 - 2 - 4 = 0



6 - 4 - 2 = 0

6 - 4 - 2 = 0

Sulfur can exceed the octet because it has empty d orbitals (it is in the 3rd energy level)

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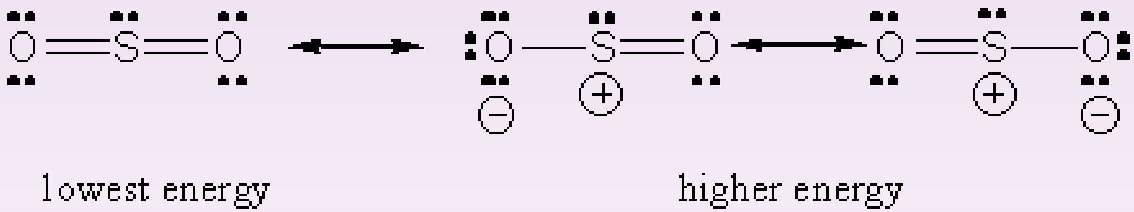
Formal charge = Valence e^- - lone pair e^- - # of bonds

DRAWING LEWIS STRUCTURES

Draw: SO_2



There are 3 resonance structures for SO_2

The preferred structure is the one where all formal charges are zero.

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DRAWING LEWIS STRUCTURES

Draw: PO_4^{3-}

O

O P O

O

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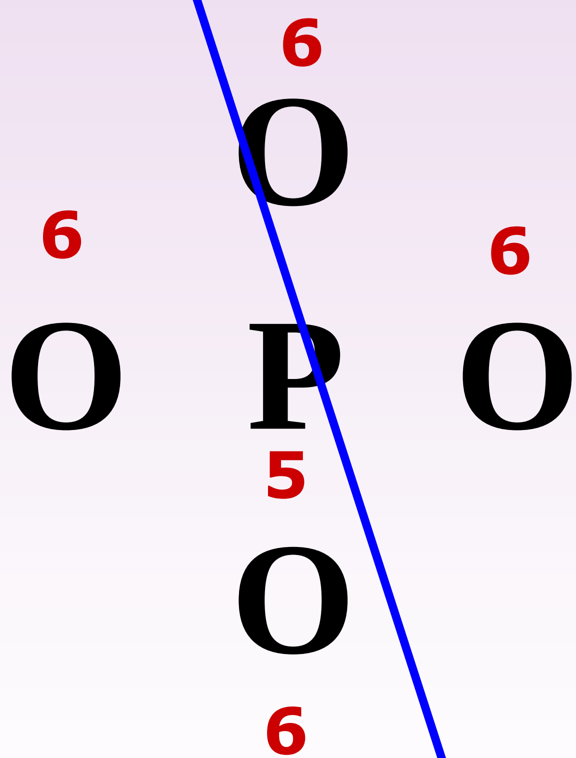
DRAWING LEWIS STRUCTURES

Draw: PO_4^{3-}



$6+6+6+6+5=29$ $+3 = 32$ electrons

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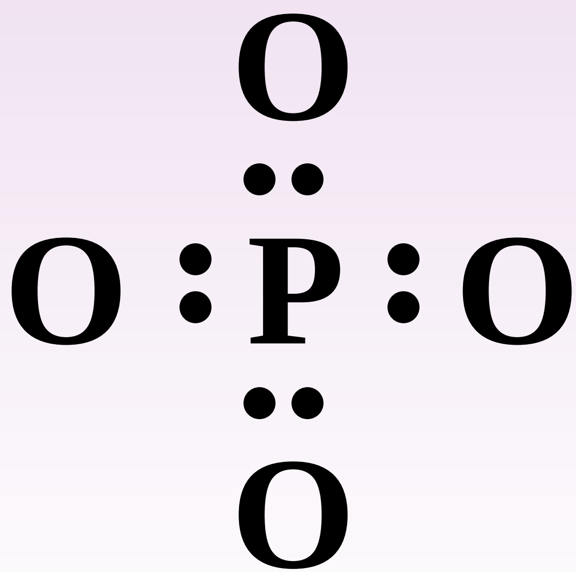
DRAWING LEWIS STRUCTURES

Draw: PO_4^{3-}



24 electrons

$6+6+6+6+5+3=$ ~~32 electrons~~

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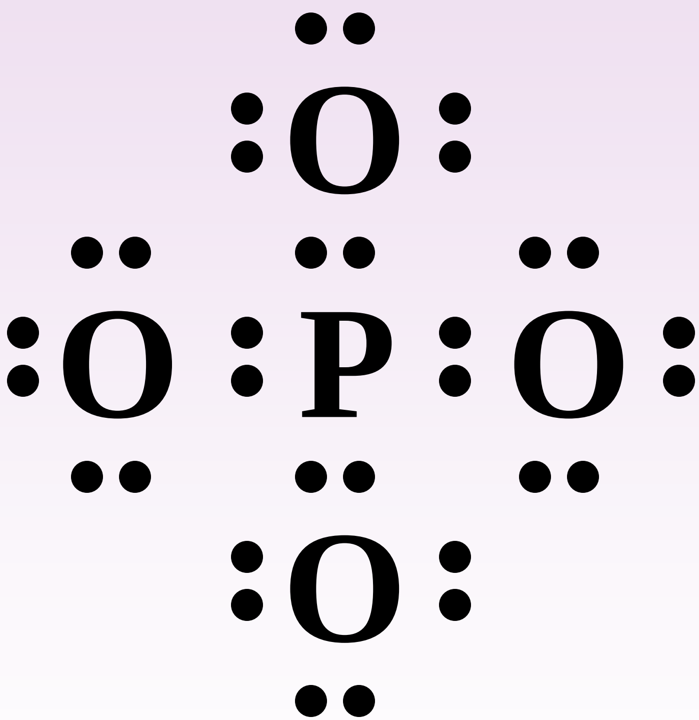
DRAWING LEWIS STRUCTURES

Draw: PO_4^{3-}



0 electrons

~~24 electrons~~ remaining

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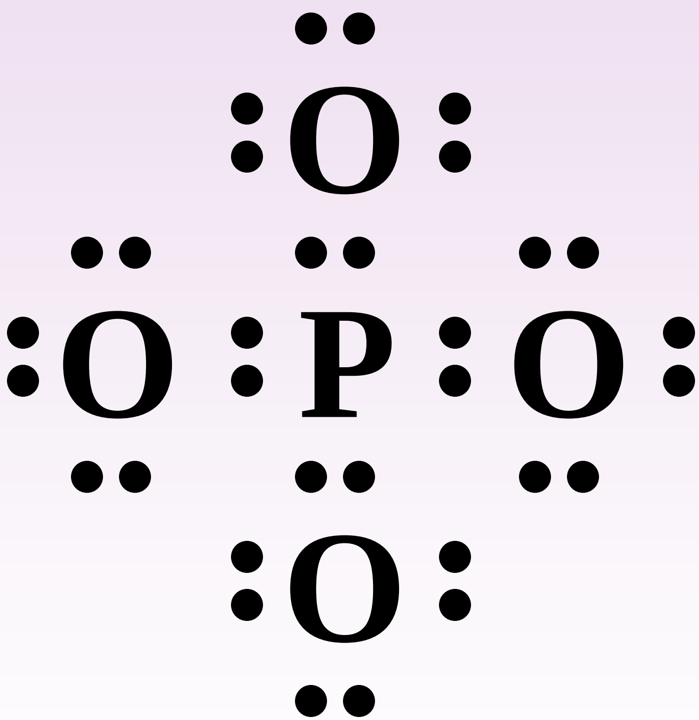
DRAWING LEWIS STRUCTURES

Draw: PO_4^{3-}



0 electrons remaining

There are no electrons left to place

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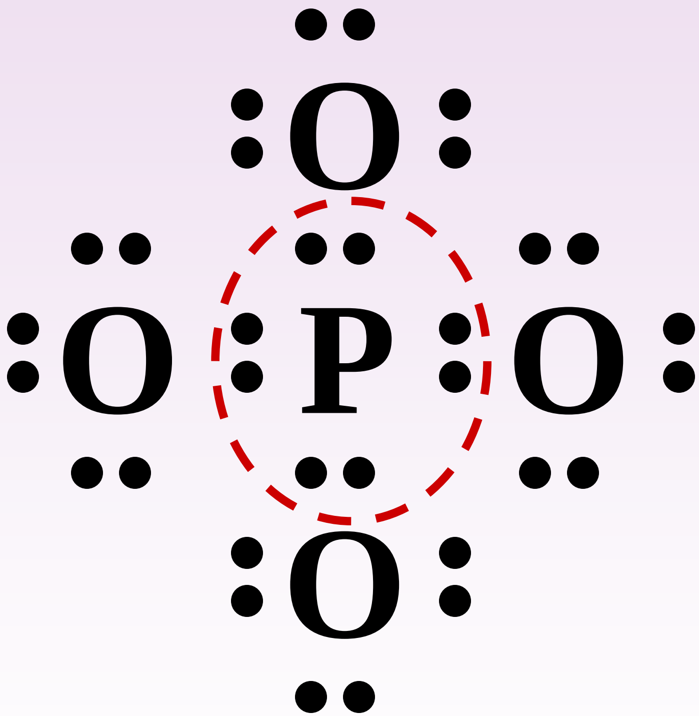
DRAWING LEWIS STRUCTURES

Draw: PO_4^{3-}



Now all atoms have a full octet

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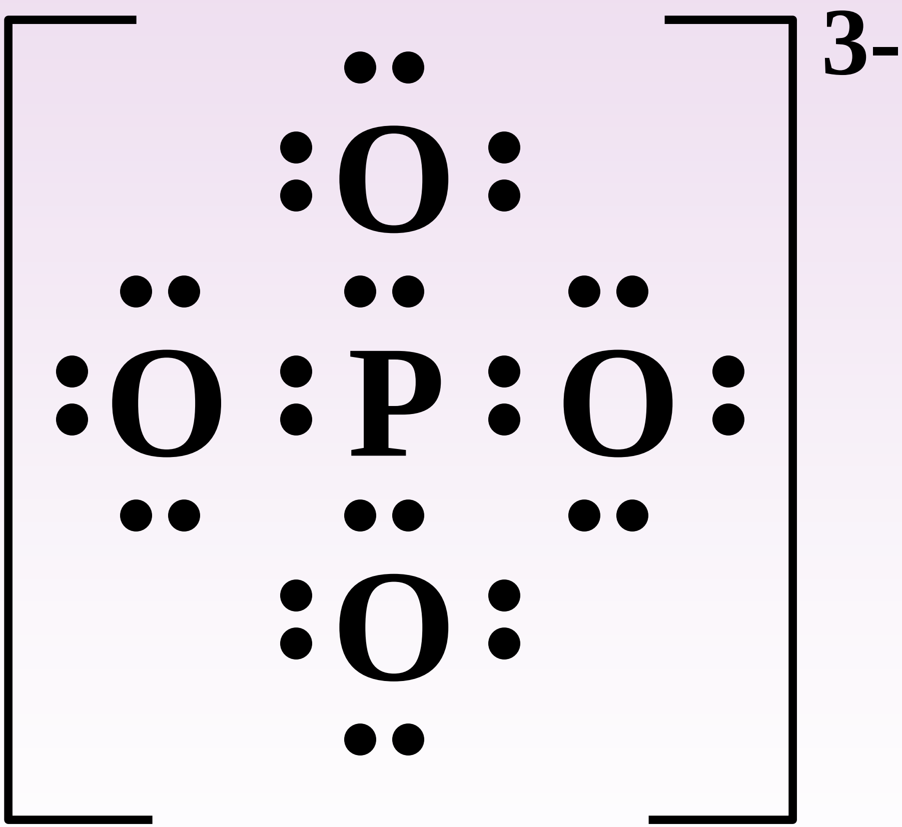
DRAWING LEWIS STRUCTURES

Draw: PO_4^{3-}



[ions must be drawn with square brackets indicating the charge]

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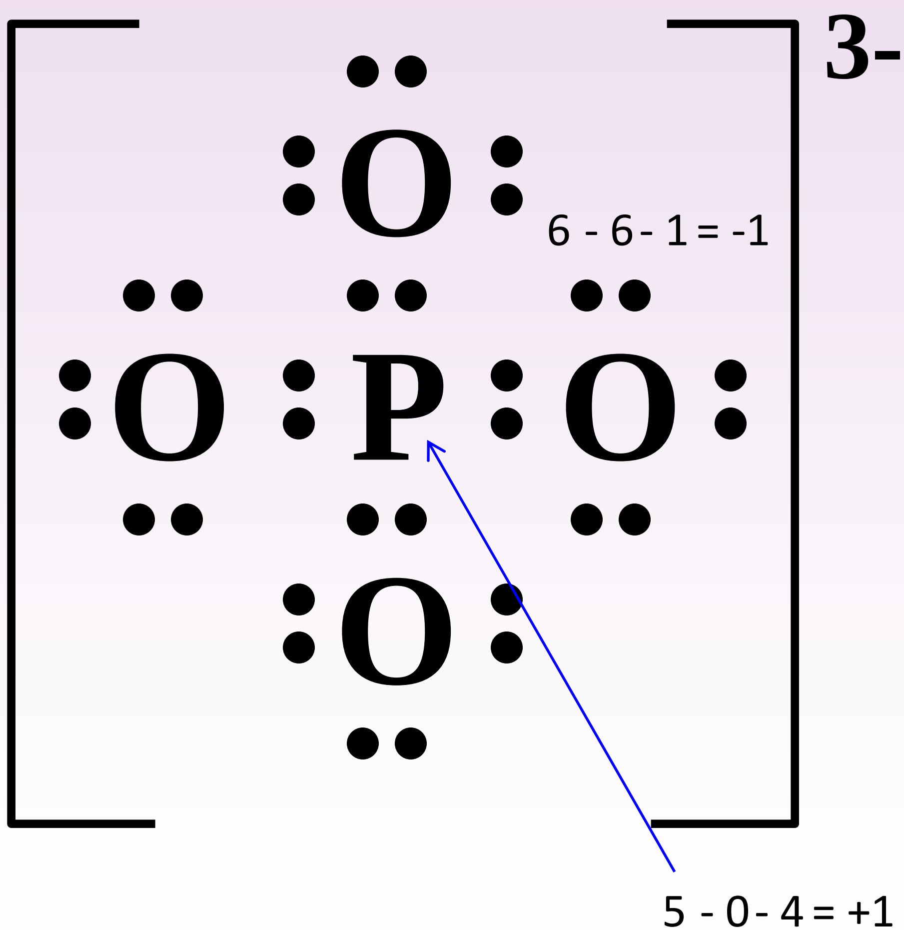
DRAWING LEWIS STRUCTURES

Draw: PO_4^{3-}



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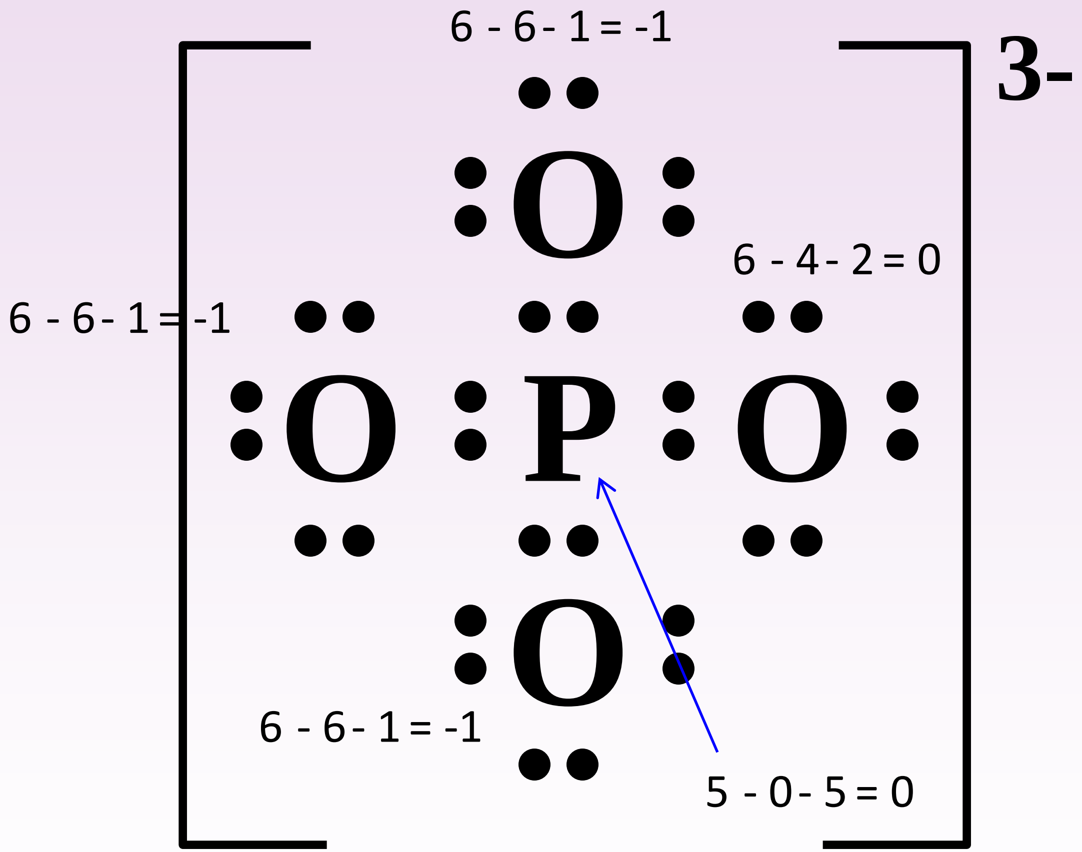
DRAWING LEWIS STRUCTURES

Draw: PO_4^{3-}



Phosphorus can exceed the octet because it has empty d orbitals
(it is in the 3^{rd} energy level)

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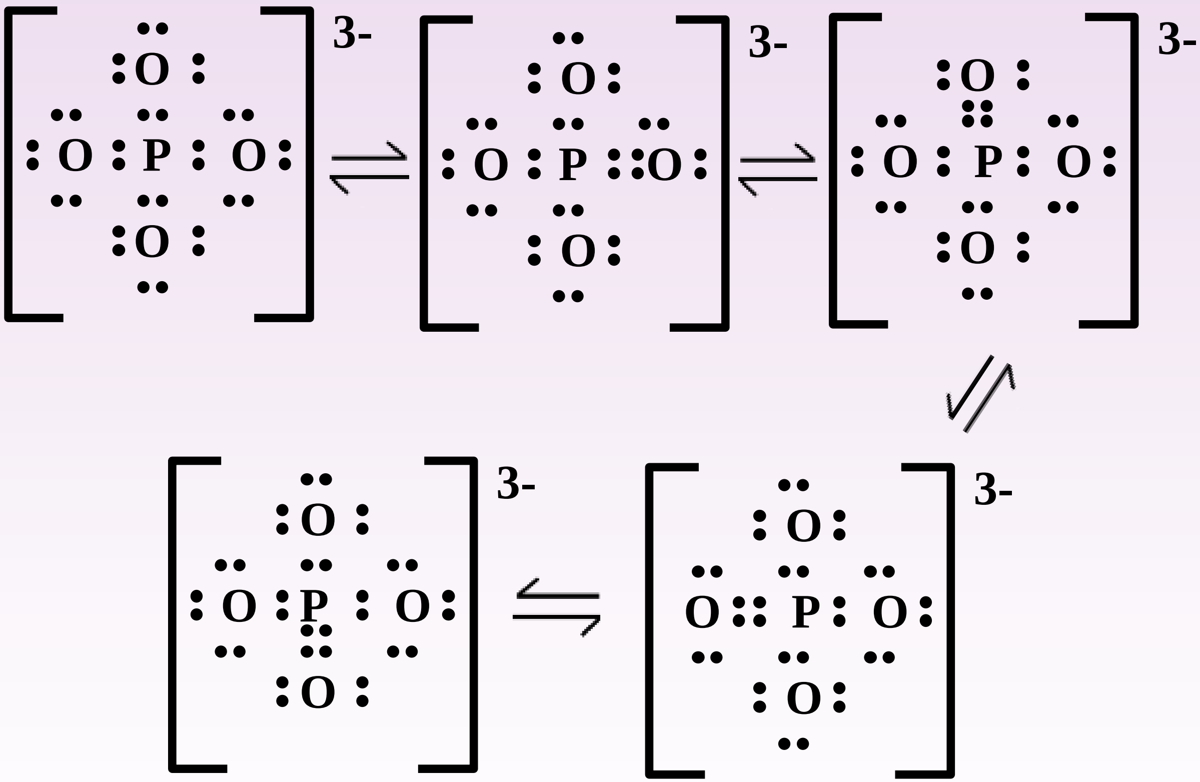
DRAWING LEWIS STRUCTURES

Draw: PO_4^{3-}



There are 5 resonance structures for PO_4^{3-}

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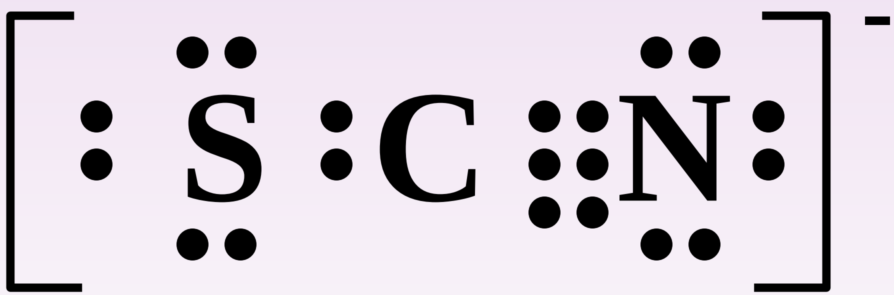
FORMAL CHARGES AND RESONANCE STRUCTURES

Draw the Lewis Dot Diagram for **SCN^-**

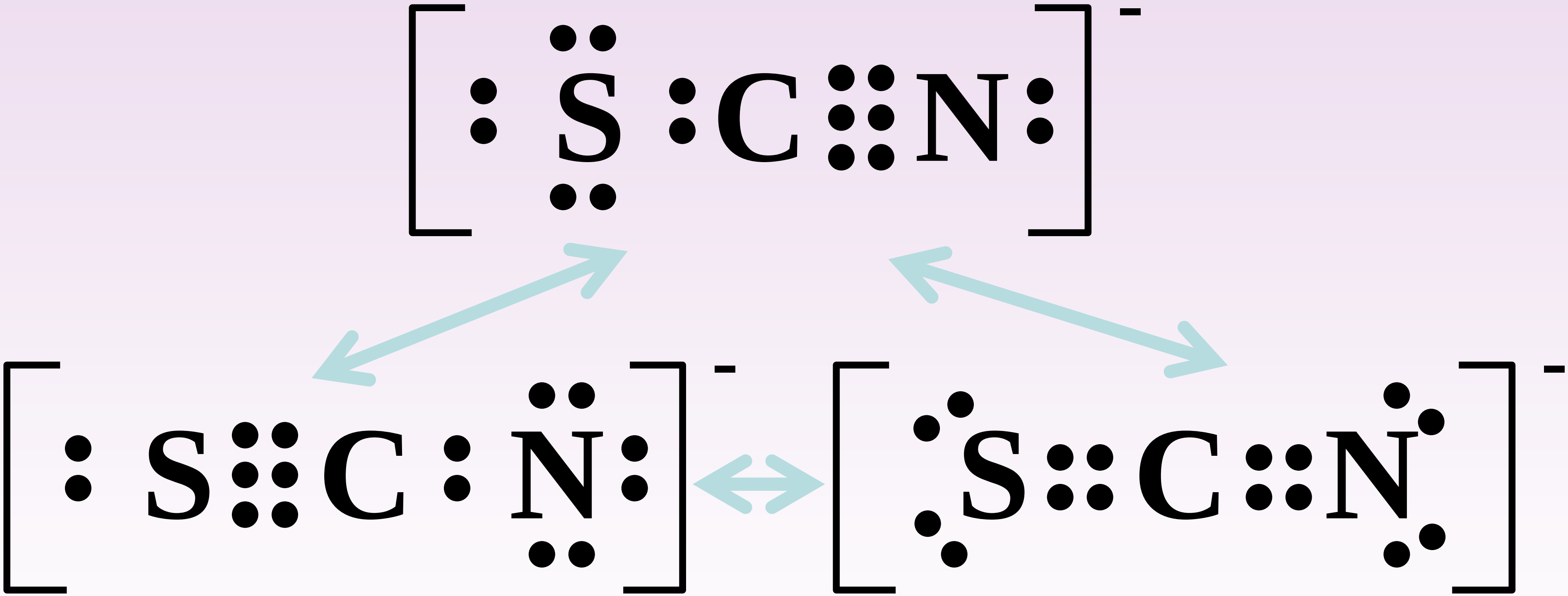


6 + 4 + 5 + 1 = 16 electrons

But this isn’t the only possible way to draw the structure

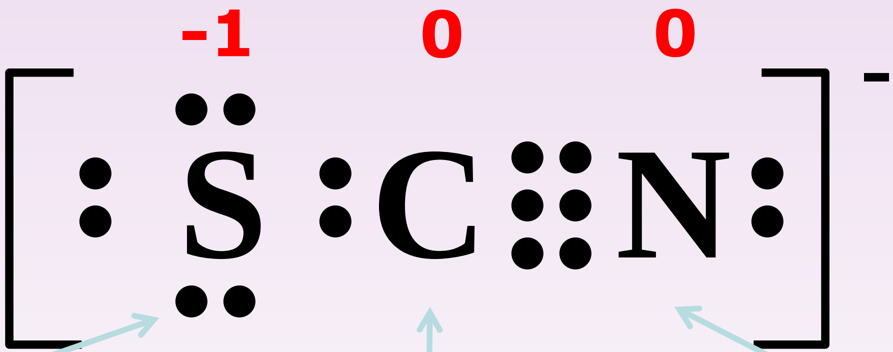
FORMAL CHARGES AND RESONANCE STRUCTURES

Draw the 3 possible Lewis Dot Diagrams for **SCN^-**



Which of these is the most likely resonance structure?

Assigning formal charges:



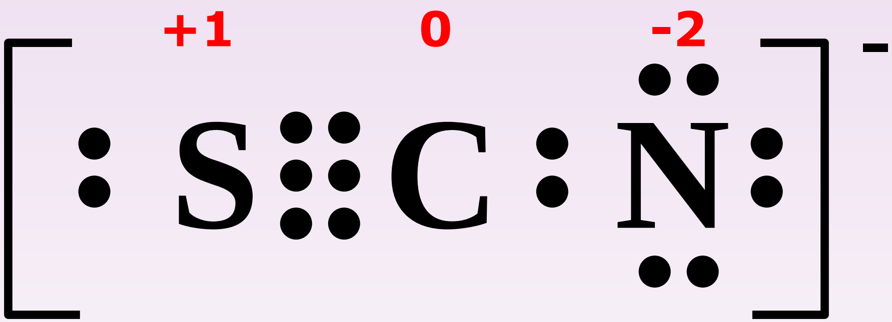
S: 6 - 6 - 1 = -1

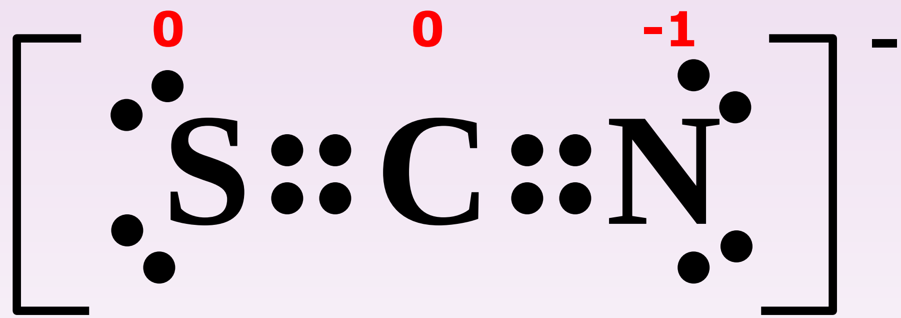
N: 5 - 2 + 3 = 0

C: 4 - 0 - 4 = 0

FORMAL CHARGES AND RESONANCE STRUCTURES

Assign formal charges to the other structures for **SCN^-**



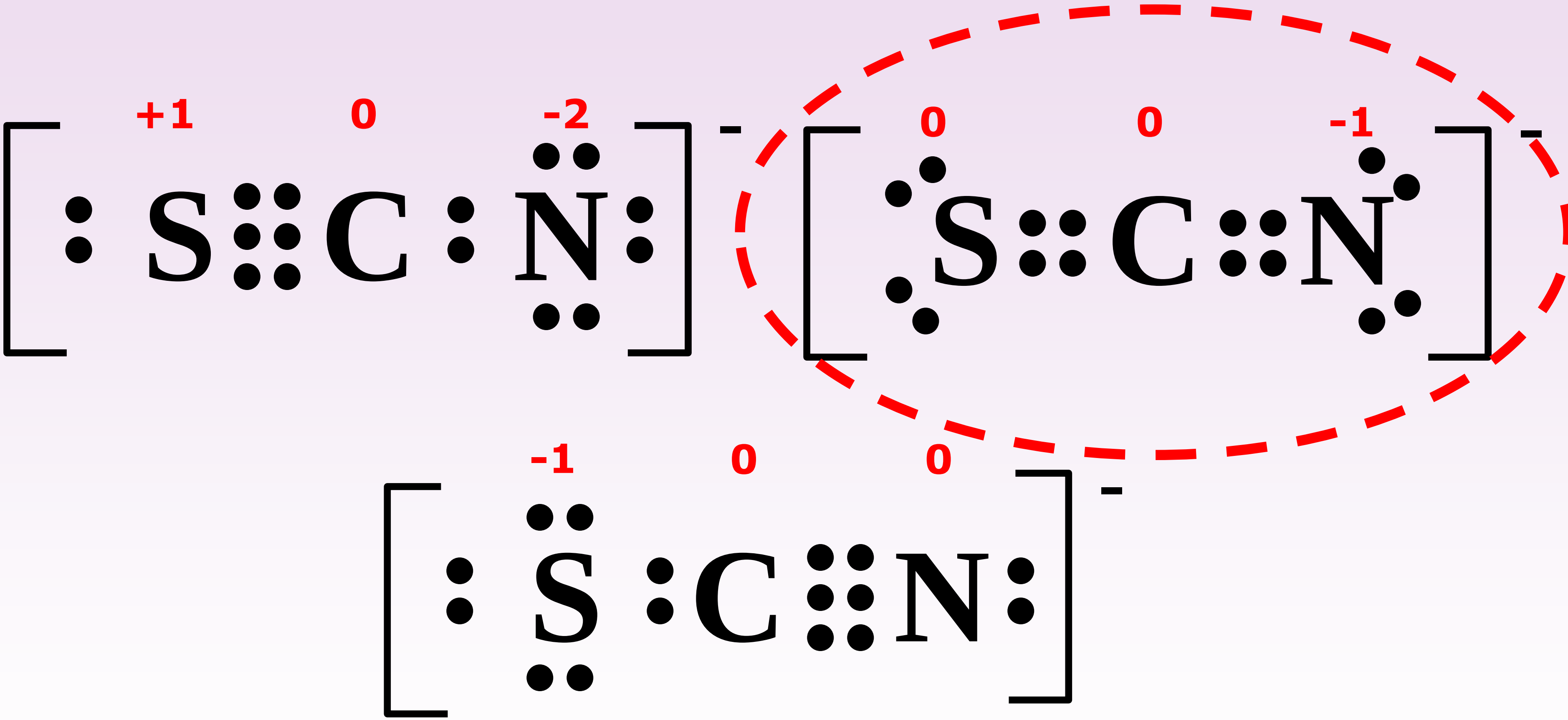


S:	**6 - 2 - 3**	**= +1**
C:	**4 - 0 - 4**	**= 0**
N:	**5 - 6 - 1**	**= -2**

S:	**6 - 4 - 2**	**= 0**
C:	**4 - 0 - 4**	**= 0**
N:	**5 - 4 - 2**	**= -1**

FORMAL CHARGES AND RESONANCE STRUCTURES

Example #1: Compare the SCN^- structures. Which one has the most neutral formal charges?

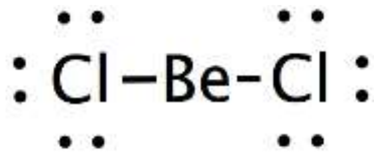


This is the most preferred structure. There are the most 0 formal charges and the most electronegative atom has the negative charge.

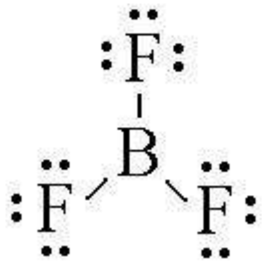
EXCEPTIONS TO THE OCTET RULE

Examples of central atoms that do not obey the octet rule:

Under-filled octets:

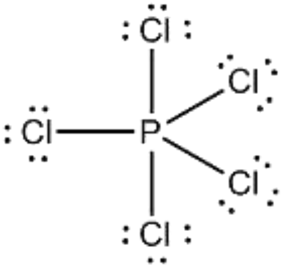


Be – 2 bonds (4 electrons)

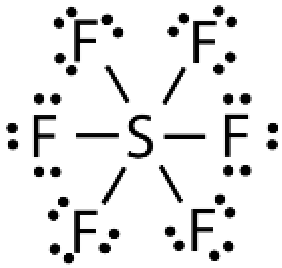


B – 3 bonds (6 electrons)

Over-filled octets:



P – 5 bonds (10 electrons)

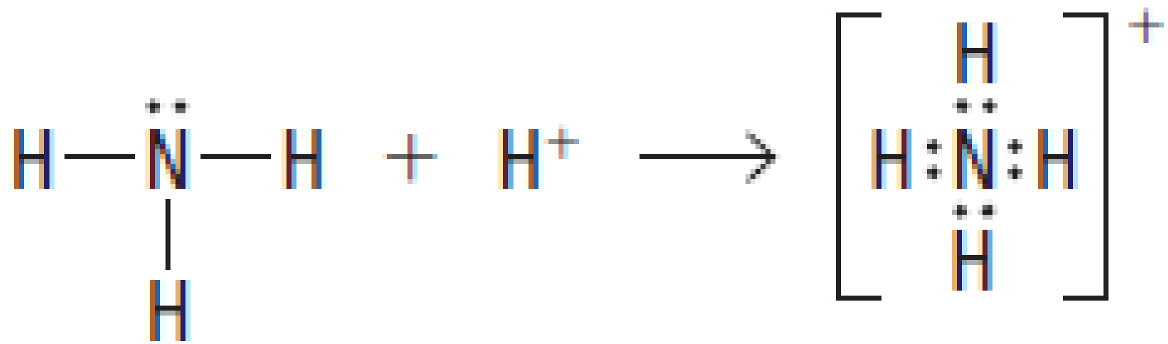


S – 6 bonds (12 electrons)

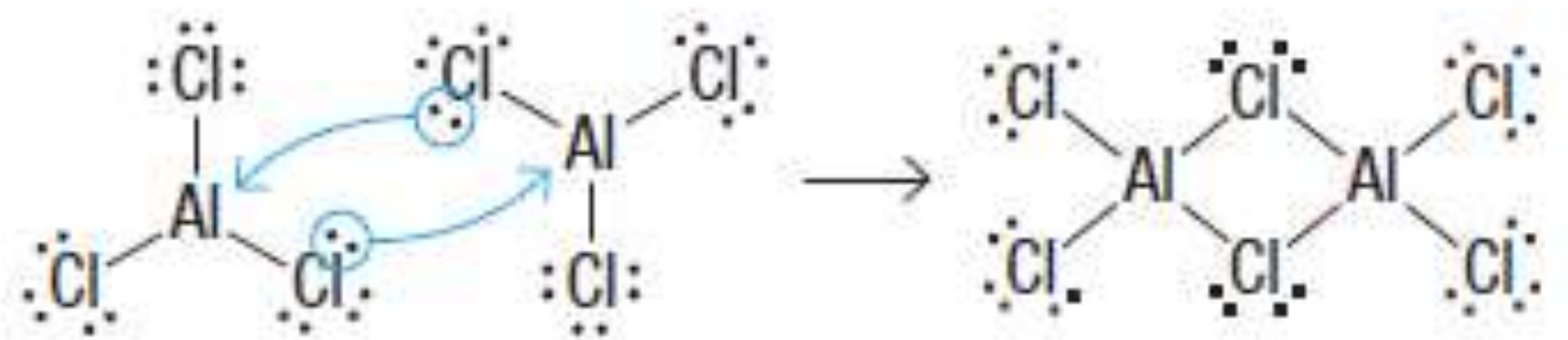
COORDINATE COVALENT BONDING

Sometimes, when a covalent bond forms between 2 atoms, both shared electrons are donated by 1 atom

Example: NH_4^+



Example: Al_2Cl_6



LEWIS STRUCTURES

Homework:

Read pages 194 & 195 (if have not already)

Page 200 #1, 2

Page 204 #1, 2

Page 205 #4, 5