Ions

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(Chemistry Notes)

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Subatomic Particles

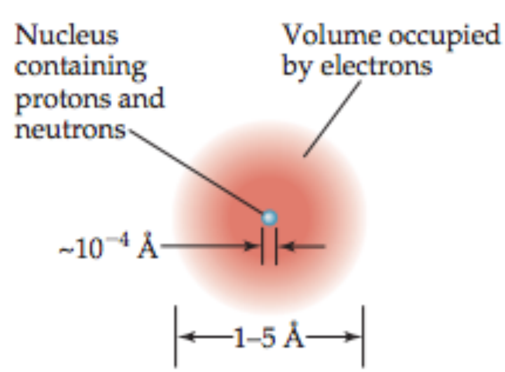
The small parts that make up atoms are called subatomic particles. Electrons (e^-) and protons (p^+) are two of the subatomic particles.

An electron has a negative charge of **-**1.602 x 10 $^{-19}$ coulombs. A proton has the same opposite charge of **+**1.602 x 10 $^{-19}$ coulombs. Protons are located in the nucleus.

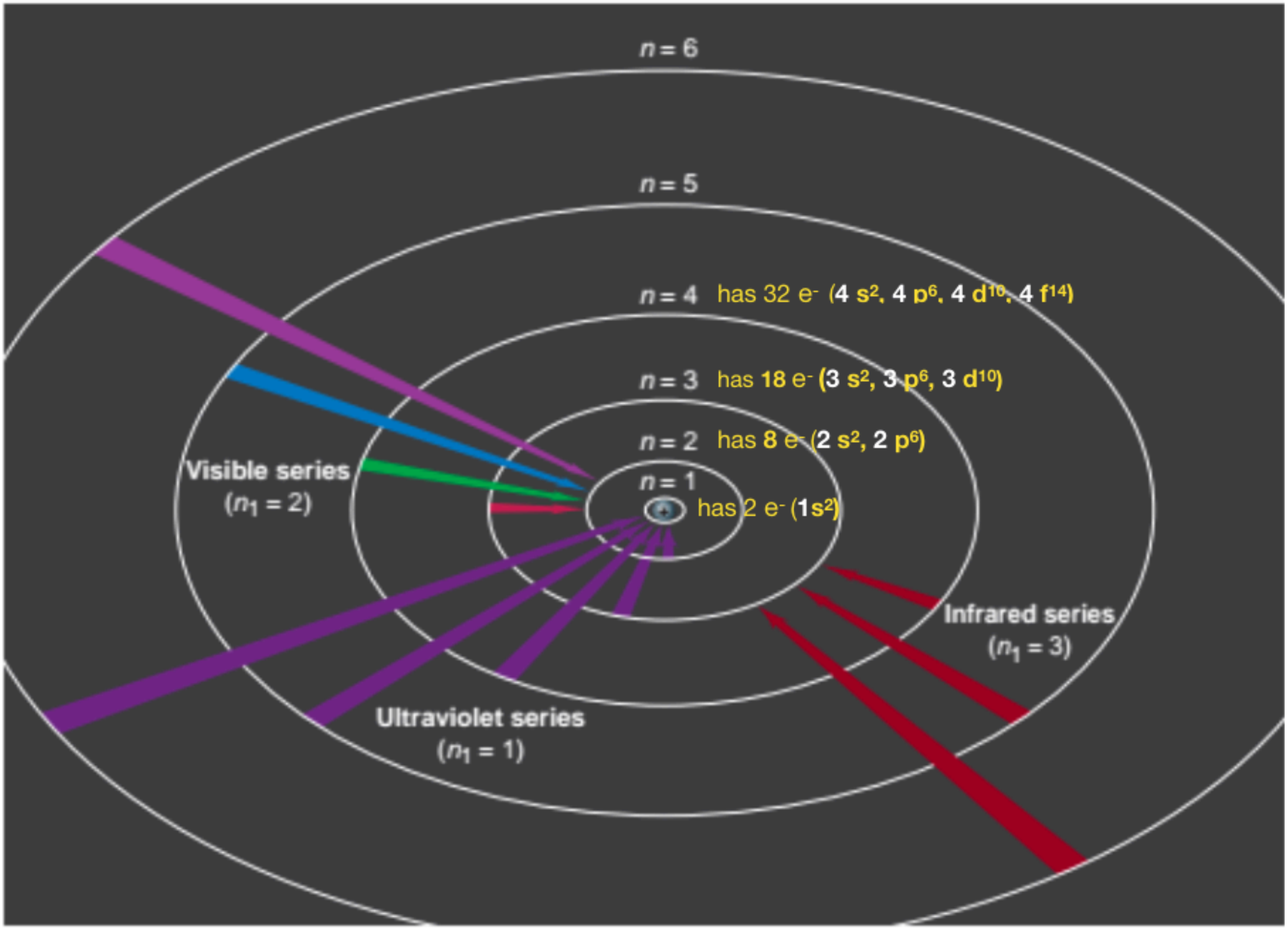
$$(1\ p^+) + (1\ e^-) = (+1.602 \times 10^{-19}) + (-1.602 \times 10^{-19}) = 0$$

Opposite charges attract each other. Therefore, there is an attraction between the nucleus and the electrons.

The nucleus is extremely small comparing to the size of the atom. The radius of the atom is 10000 times larger than the radius of the nucleus. If the nucleus of an atom were the size of a marble, then the whole atom would be about the size of a football stadium.



The simplest representation of an atom:



Ions

An ion is a charged atom. The charge can be positive (+) or negative (-). There are two types of ions: Cations and Anions.

Cations are positively charged ions (X^{n+}). Example: Na^{+}

Anions are negatively charged ions (X^{n-}). Example: Cl^{-}

Why Do Atoms Form Ions?

The Noble Gases (Group 8A or 18)

Noble gases are **stable**. They are the elements of group 8A or 18. They have 8 valence electrons (8 electrons on the outer energy level ($n\,S^{2}$, $n\,P^{6}$), except He. They are: He, Ne, Ar, Kr, Xe and Rn. Another name for the number **8 is octet**.

Octet Rule:

In most chemical reactions, **atoms** tend to match the S and P electron configurations of the noble gases so they can become more stable. In other words, atoms tend to have 8 electrons on the outer shell **similar to the noble gases**. This tendency is called the octet rule.

See table in the next page.

<table>
<tr><th>Helium (He)
Z = 2</th><th>Neon (Ne)
Z = 10</th><th>Argon (Ar)
Z = 18</th><th>Krypton (Kr)
Z = 36</th></tr>
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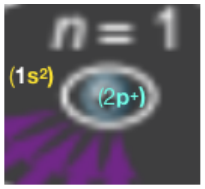
<td>
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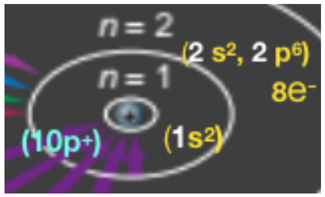
<td>
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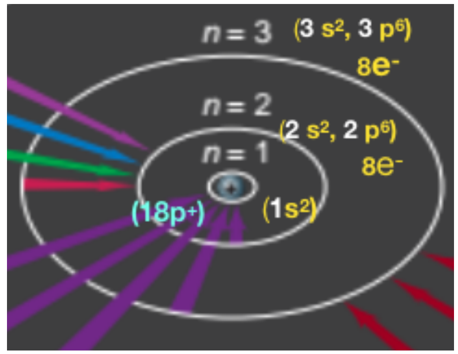
<td>
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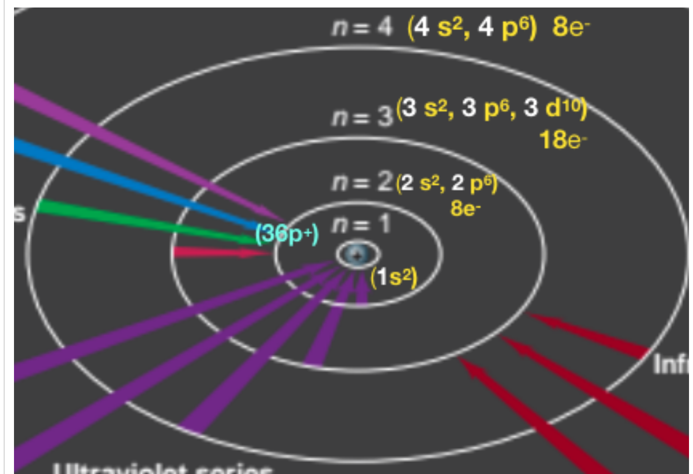
<td>
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</tr>
<tr><td>$1s^2$</td><td>$1s^2$, $2s^2$, $2p^6$</td><td>$1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^6$</td><td>$1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^6$, $4s^2$, $3d^{10}$, $4s^2$, $4p^6$</td></tr>
<tr><td>1 Energy level
(n = 1)

Period 1</td><td>2 Energy levels
(n = 2)

Period 2</td><td>3 Energy levels
(n = 3)

Period 3</td><td>4 Energy levels
(n = 4)

Period 4</td></tr>
<tr><td colspan="4">All have 8 electrons on the outer shell (8 valence electrons) (s^2, p^6)

Group 8A or 18</td></tr>
</table>

Cations and Anions

An ion is a charged atom. The charge can be positive (+) or negative (-). There are two types of ions: Cations and Anions. Cations are ions with a positive charge (X^{n+}). Anions are ions with a negative charge (X^{n-}). Below is a summary. Explanations are in the following pages.

Cation X^{n+}	When an atom (X) loses n electrons, it becomes positively charged ion or cation (X^{n+}). **Equation:** $X - n\,e^- = X^{n+}$ **Size:** The size of the cation is smaller than the size of the parent atom. $X^{n+} < X$ **Name**: The name of the cation is the same as that of the parent element: Na^+ is called sodium cation.
Anion Y^{n-}	When an atom (Y) gains n electrons, it becomes negatively charged ion or anion (Y^{n-}). **Equation:** $Y + n\,e^- = Y^{n-}$ **Size:** The size of the anion is smaller than the size of the parent atom. $Y^{n-} > Y$ **Name**: The name of the anion is formed from the name of the element, but it ends with ***-ide***. Cl^- is called chlor**ide**.

Ions of Elements of Group 1A: Summary (explanation is on next page)

Each element of group 1 loses the one valence electron (e^-) to look like one the noble gases (usually the closest to it) and becomes more stable. When an atom loses one electron, it becomes positively charged ion or cation (X^+). X^+ has the electron configuration like that of a noble gas.

$$X - 1\ e^- = X^+$$

Example:

Potassium (K), in group 1, has an atomic number (Z) = 19. Potassium has 19 electrons and 19 protons. The noble gas that has the closest atomic number is Argon (Ar, Z = 18). So, Ar has 18 electrons.

To be more stable, potassium (19 electrons) has to lose one electron so it has the same number of electrons as Ar (18 electrons).

Potassium can lose the one valence electron on the outer shell. Then, it will have only 18 electrons (18 e^-) left. The number of protons in the nucleus stays 19 protons (19 p^+).

The new charge on potassium will be: $(19^+) + (18^-) = 1^+$ or K^+

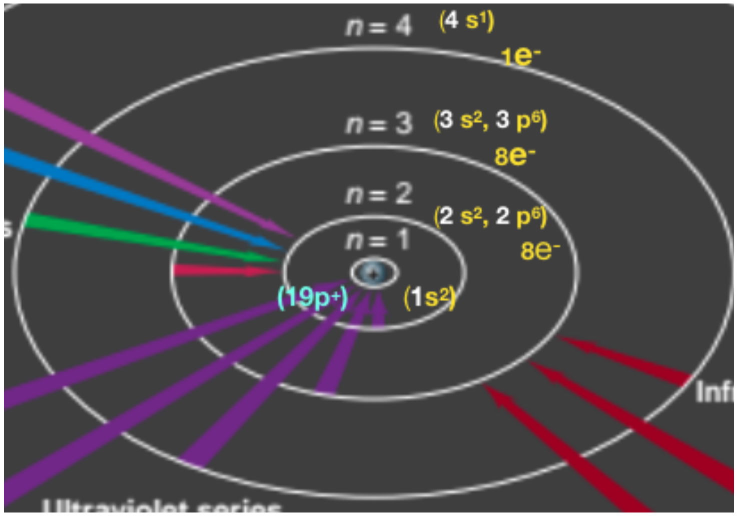
Equation: $K - 1\ e^- = K^+$

Size: The size of the cation is smaller than the size of the parent atom.

$$K^+ < K$$

Name: The name of the cation is the same as that of the parent atom: K^+ is called potassium cation.

Both K^+ and Ar have the same electron configurations.



Potassium (K)

Z = 19
19 P^+
<u>19 e^-</u>
Total Charge = 0
Neutral atom

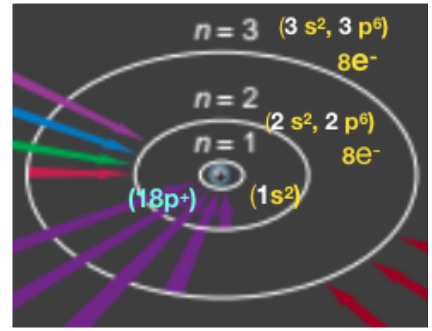
4 energy levels (period 4)
1 valence electron (group 1A)

1 s^2, 2 s^2, 2 p^6 , 3 s^2, 3 p^6, 4 s^1

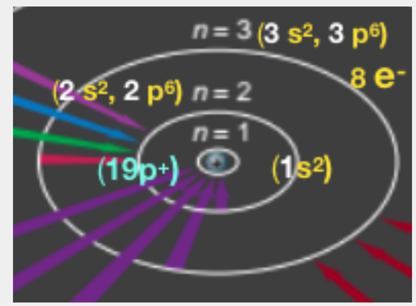
- The closest noble gas to potassium (Z = 19) is Argon (Z = 18).
- Potassium can be more stable by loosing the 1 valence electron (4S¹).
- Then it will look like argon with 8 valence electrons (Octet).
- The nucleus does not change (same number of protons).

$K - 1\ e^- = K^+$

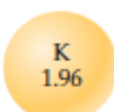
- K^+ and Ar have the same electron configuration.



Argon (Ar, Z = 18)
8 valence electrons (Octet, stable)
18 P^+, 18 e^-
1 s^2, 2 s^2, 2 p^6 , 3 s^2, 3 p^6



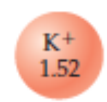
K^+
8 valence electrons (Octet, stable)
19 P^+
<u>18 e^-</u>
Total Charge = 1+
3 energy levels
1 s^2, 2 s^2, 2 p^6 , 3 s^2, 3 p^6



Potassium radii = 1.96 angstroms

K^+ has 3 energy levels.
K has 4 energy levels.

K^+ is smaller than K



Potassium cation radii = 1.52 angstroms

Ions of Elements of Group 2A: Summary (explanation is on next page)

Each element of group 2 loses the two valence electron (e^-) to look like one the noble gases (usually the closest to it) and becomes more stable. When an atom loses 2 electron, it becomes positively charged ion or cation (X^{2+}). X^{2+} has the electron configuration like that of a noble gas.

$$X - 2\,e^- = X^{2+}$$

Example:

Magnesium (Mg), in group 2, has an atomic number (Z) = 12. Potassium has 12 electrons and 12 protons.
The noble gas that has the closest atomic number is Ne (Ne, Z = 10). So, Ne has 10 electrons.

To be more stable, Magnesium (12 electrons) has to lose 2 electrons so it has the same number of electrons as Ne (10 electrons).

Magnesium can lose the two valence electron on the outer shell. Then, it will have only 10 electrons ($10\,e^-$) left. The number of protons in the nucleus stays 12 protons ($12\,p^+$).

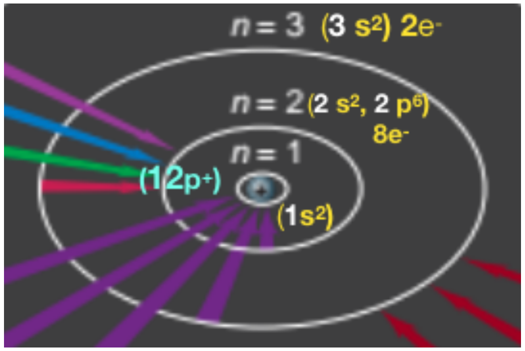
The new charge on magnesium will be: $(12^+) + (10^-) = 2^+$ or Mg^{2+}

Equation: $$Mg - 2\,e^- = Mg^{2+}$$

Size: The size of the magnesium cation is smaller than the size of the parent atom.

$$Mg^{2+} < Mg$$

Name: Mg^{2+} is called magnesium cation.



Magnesium (Mg)

Z = 12
12 P+
<u>12 e-</u>
Total Charge = 0
Neutral atom

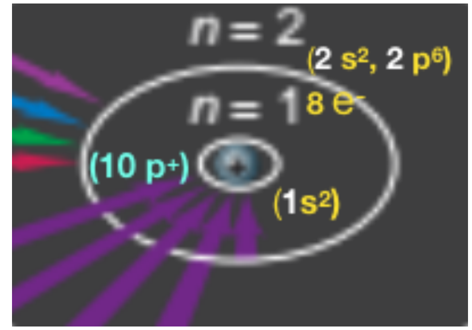
3 energy levels (period 3)
2 valence electron (group 2A)

1 **s²**, 2 **s²**, 2 **p⁶** , 3 **s²**

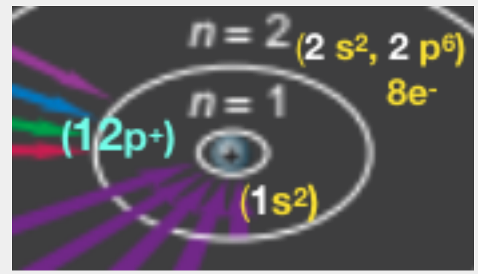
- The closest noble gas to Magnesium (Z = 12) is Neon (Z = 10).
- Magnesium can be more stable by loosing the 2 valence electron (3S²).
- Then it will look like argon with 8 valence electrons (Octet).
- The nucleus does not change (same number of protons).

$Mg - 2e^{-} = Mg^{2+}$

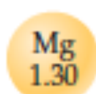
- Mg^{2+} and Ne have the same electron configuration.



Neon (Ne, Z = 10)
8 valence electrons (Octet, stable)
10 P+, 10 e-
1 **s²**, 2 **s²**, 2 **p⁶**



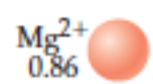
Mg^{2+}
8 valence electrons (Octet, stable)
12 P+
<u>10 e-</u>
Total Charge = 2+
2 energy levels
1 **s²**, 2 **s²**, 2 **p⁶**



Magnesium radii = 1.30 angstroms

Mg^{2+} has 2 energy levels.
Mg has 3 energy levels.

Mg^{2+} is smaller than Mg



Magnesium cation radii = 0.86 angstroms

Ions of Elements of Group 3A or 13: Summary (explanation is on next page)

An element of group 3 can lose the three valence electron (e^-) to look like one the noble gases (usually the closest to it) and becomes more stable. When an atom loses 3 electron, it becomes positively charged ion or cation (X^{3+}). X^{3+} has the electron configuration like that of a noble gas.

$$X - 3\ e^- = X^{3+}$$

Example:

Aluminum (Al), in group 3, has an atomic number (Z) = 13. Aluminum has 13 electrons and 13 protons. The noble gas that has the closest atomic number is Ne (Ne, Z = 10). So, Ne has 10 electrons.

To be more stable, Aluminum (13 electrons) has to lose 3 electrons so it has the same number of electrons as Ne (10 electrons).

Aluminum can lose the three valence electron on the outer shell. Then, it will have only 10 electrons (10 e^-) left. The number of protons in the nucleus stays 13 protons (13 p^+).

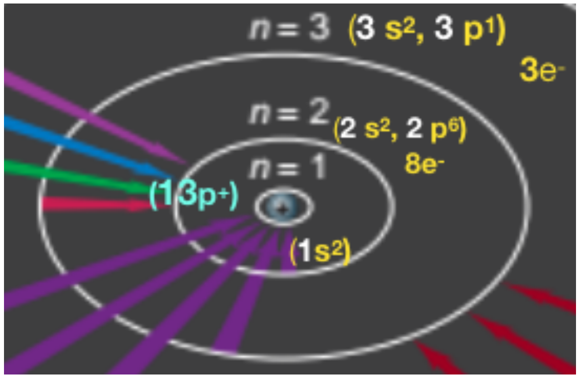
The new charge on Aluminum will be: $(13^+) + (10^-) = 3^+$ or Al^{3+}

Equation: $Al - 3\ e^- = Al^{3+}$

Size: The size of the cation is smaller than the size of the parent atom.

$$Al^{3+} < Al$$

Name: Al^{3+} is called aluminum cation.



Aluminium (Al)

Z = 13
13 P^+
<u>13 e^-</u>
Total Charge = 0
Neutral atom

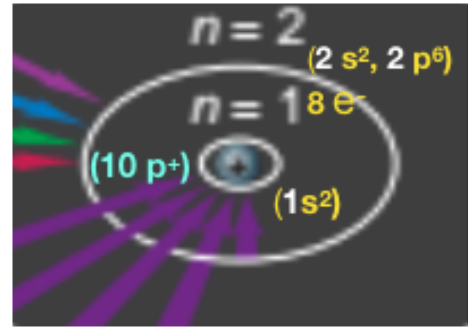
3 energy levels (period 3)
3 valence electron (group 3 A or 13)

1 **s^2**, 2 **s^2**, 2 **p^6** , 3 **s^2**, 3 **p^1**

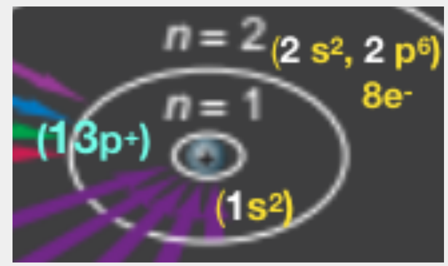
- The closest noble gas to Aluminum (Z = 13) is Neon (Z = 10).
- Aluminum can be more stable by loosing the 3 valence electrons ($3S^2$, $3P^1$).
- Then it will look like neon with 8 valence electrons.
- The nucleus does not change (same number of protons).

AL - 3 e^- = Al $^{3+}$

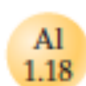
- Al^{3+} and Ne have the same electron configuration.



Neon (Ne, Z = 10)
8 valence electrons (Octet, stable)
10 P^+, 10 e^-
1 **s^2**, 2 **s^2**, 2 **p^6**



Al^{3+}
8 valence electrons (Octet, stable)
13 P^+
<u>10 e^-</u>
Total Charge = 3+
2 energy levels
1 **s^2**, 2 **s^2**, 2 **p^6**



Aluminum radii = 1.18 angstroms

Al^{3+} has **2** energy levels.
Al has **3** energy levels.

Al^{3+} is smaller than Al

Aluminum cation radii = 0.68 angstroms

Ions of Elements of Group 5A or 15: Summary (explanation is on next page)

The elements of group 5 have 5 valence electron (e^-). To look like one the noble gases with 8 valence electron, the element can gain three more electrons . When an atom gains 3 electrons, it becomes negatively charged ion or anion (X^{3-}). X^{3-} has the electron configuration like that of a noble gas.

$$X + 3\,e^- = X^{3-}$$

Example:

Nitrogen (N), in group 3, has an atomic number (Z) = 7. Nitrogen has 7 electrons and 7 protons.
The noble gas that has the closest atomic number is Ne (Ne, Z = 10). So, Ne has 10 electrons.

To be more stable, nitrogen (7 electrons) has to gain 3 electrons so it has the same number of electrons as Ne (10 electrons). Then, it will have only 10 electrons (10 e^-). The number of protons in the nucleus stays 7 protons (7 p^+).

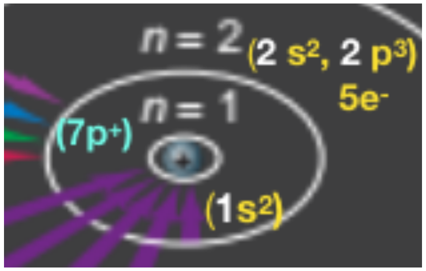
The new charge on nitrogen will be: $(7^+) + (10^-) = 3^-$ or N^{3-}

Equation: $N + 3\,e^- = N^{3-}$

Size: The size of the anion is bigger than the size of the parent atom.

$$N^{3-} < N$$

Name: N^{3-} is called nitr**ide**.



Nitrogen (N)

Z = 7
7 P^+
<u>7 e^-</u>
Total Charge = 0
Neutral atom

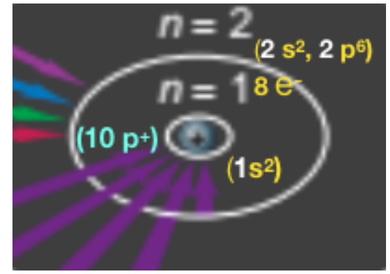
2 energy levels (period 2)
6 valence electron (group 5 A or 15)

1 $\mathbf{s^2}$, 2 $\mathbf{s^2}$, 2 $\mathbf{p^3}$

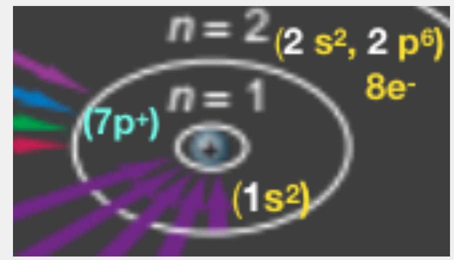
- The closest noble gas to nitrogen(Z = 7) is Neon (Z = 10).
- nitrogen can be more stable by gaining the 3 more valence electrons ($2P^{3+3}$**)**.
- Then it will look like neon with 8 valence electrons.
- The nucleus does not change (same number of protons).

$$N + 3\ e^- = N^{3-}$$

- $\mathbf{N^{3-}}$ and Ne have the same electron configuration.



Neon (Ne, Z = 10)
8 valence electrons (Octet, stable)
10 P^+, 10 e^-
1 $\mathbf{s^2}$, 2 $\mathbf{s^2}$, 2 $\mathbf{p^6}$



N^{3-}
8 valence electrons (Octet, stable)
7 P^+
<u>10 e^-</u>
Total Charge = 3-
2 energy levels
1 $\mathbf{s^2}$, 2 $\mathbf{s^2}$, 2 $\mathbf{p^6}$

Both **N** and $\mathbf{N^{3-}}$ have 2 energy levels.
$\mathbf{N^{3-}}$: 7 P^+ are pulling 10 e^-
N : 7 P^+ can pull its 7 e^- closer.
N is smaller than N^{3-}

Ions of Elements of Group 6A or 16: Summary (explanation is on next page)

The elements of group 6 have 6 valence electron (e^-). To look like one the noble gases with 8 valence electron, the element can gain two more electrons . When an atom gains 2 electrons, it becomes negatively charged ion or anion (X^{2-}). X^{2-} has the electron configuration like that of a noble gas.

$$X + 2\ e^- = X^{2-}$$

Example:

Oxygen (O), in group 6A, has an atomic number (Z) = 8. Oxygen has 8 electrons and 8 protons.
The noble gas that has the closest atomic number is Ne (Ne, Z = 10). So, Ne has 10 electrons.

To be more stable, oxygen (8 electrons) has to gain 2 electrons so it has the same number of electrons as Ne (10 electrons). Then, it will have only 10 electrons (10 e^-). The number of protons in the nucleus stays 8 protons (8 p^+).

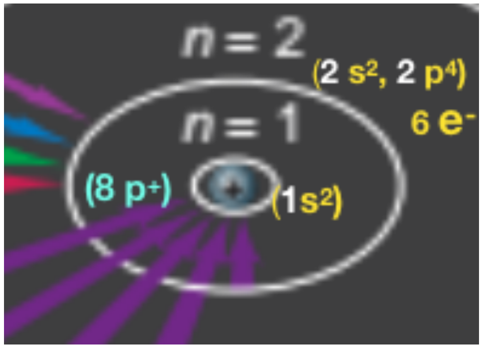
The new charge on oxygen will be: $(8^+) + (10^-) = 2^-$ or O^{2-}

Equation: $$O + 2\ e^- = O^{2-}$$

Size: The size of the anion is bigger than the size of the parent atom oxygen.

$$O^{2-} < O$$

Name: O^{2-} is called ox**ide**.



Oxygen (O)

Z = 8
8 P^+
<u>8 e^-</u>
Total Charge = 0
Neutral atom

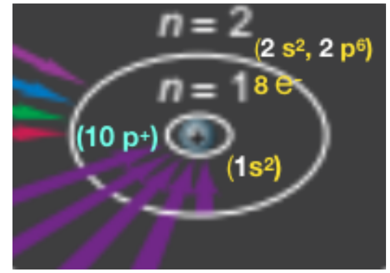
2 energy levels (period 2)
6 valence electron (group 6 A or 16)

1 s^2, 2 s^2, 2 p^4

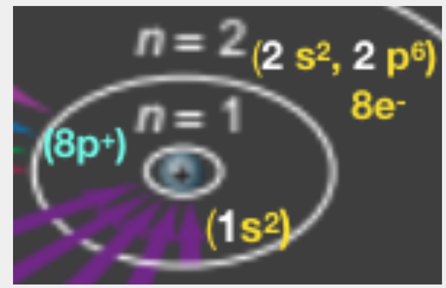
- The closest noble gas to oxygen (Z = 8) is Neon (Z = 10).
- Oxygen can be more stable by gaining the 2 more valence electrons ($2P^{4+2}$).
- Then it will look like neon with 8 valence electrons.
- The nucleus does not change (same number of protons).

$$O + 2e^- = O^{2-}$$

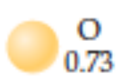
- O^{2-} and Ne have the same electron configuration.



Neon (Ne, Z = 10)
8 valence electrons (Octet, stable)
10 P^+, 10 e^-
1 s^2, 2 s^2, 2 p^6

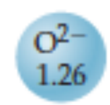


O^{2-}
8 valence electrons (Octet, stable)
8 P^+
<u>10 e^-</u>
Total Charge = 2-
2 energy levels
1 s^2, 2 s^2, 2 p^6



Oxygen radii = 0.73 angstroms

Both **O** and **O^{2-}** have **2** energy levels.
O^{2-} : 8 P^+ are pulling 10 e^-
O :8 P^+ can pull its 8 e^- closer.
O is smaller than O^{2-}



oxide radii = 1.26 angstroms

Ions of Elements of Group 7A or 17: Summary (explanation is on next page)

The elements of group 7 have 7 valence electron (e^-). To look like one the noble gases with 8 valence electron, the element can gain one more electrons . When an atom gains 1 electron, it becomes negatively charged ion or anion (X^-). X^- has the electron configuration like that of a noble gas.

$$X + \mathbf{1}\ e^- = X^-$$

Example:

Fluorine (F), in group 7A, has an atomic number (Z) = 9. Oxygen has 9 electrons and 9 protons.
The noble gas that has the closest atomic number is Ne (Ne, Z = 10). So, Ne has 10 electrons.

To be more stable, fluorine (9 electrons) has to gain 1 electron so it has the same number of electrons as Ne (10 electrons). Then, it will have only 10 electrons (10 e^-). The number of protons in the nucleus stays 9 protons (9 p^+).

The new charge on fluorine will be: $(9^+) + (10^-) = 1^-$ or F^-

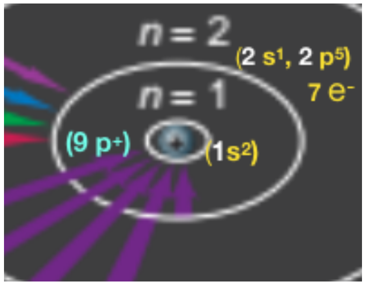
Equation: $F + 1\ e^- = F^-$

Size: The size of the anion is bigger than the size of the parent atom oxygen.

$$F^- < F$$

Name: F^- is called fluor**ide**.

Both F^- and neon have the same electron configurations.



Fluorine (F)

Z = 9

9 P^+

9 e^-

Total Charge = 0

Neutral atom

2 energy levels (period 2)

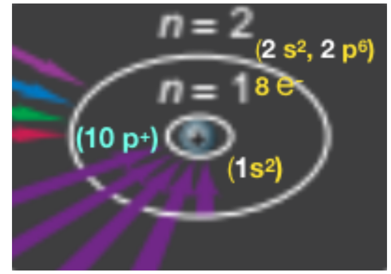
7 valence electron (group 7 A or 17)

1 s^2, 2 s^2, 2 p^5

- The closest noble gas to fluorine (Z = 9) is Neon (Z = 10).
- Fluorine can be more stable by gaining 1 more valence electrons ($2P^{5+1}$).
- Then it will look like neon with 8 valence electrons.
- The nucleus does not change (same number of protons).

$F + 1\ e^- = F^-$

- F^{2-} and Ne have the same electron configuration.

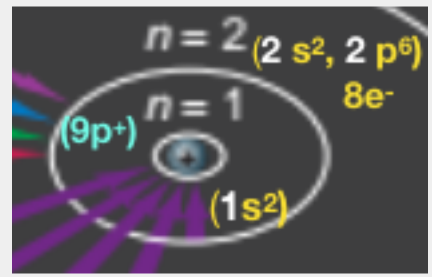


Neon (Ne, Z = 10)

8 valence electrons (Octet, stable)

10 P^+, 10 e^-

1 s^2, 2 s^2, 2 p^6



F^-

8 valence electrons (Octet, stable)

9 P^+

10 e^-

Total Charge = 1-

2 energy levels

1 s^2, 2 s^2, 2 p^6

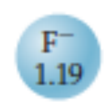
Fluorine radii =0.71 angstroms

Both **F** and **F^-** have **2** energy levels.

F^- : 9 P^+ are pulling 10 e^-

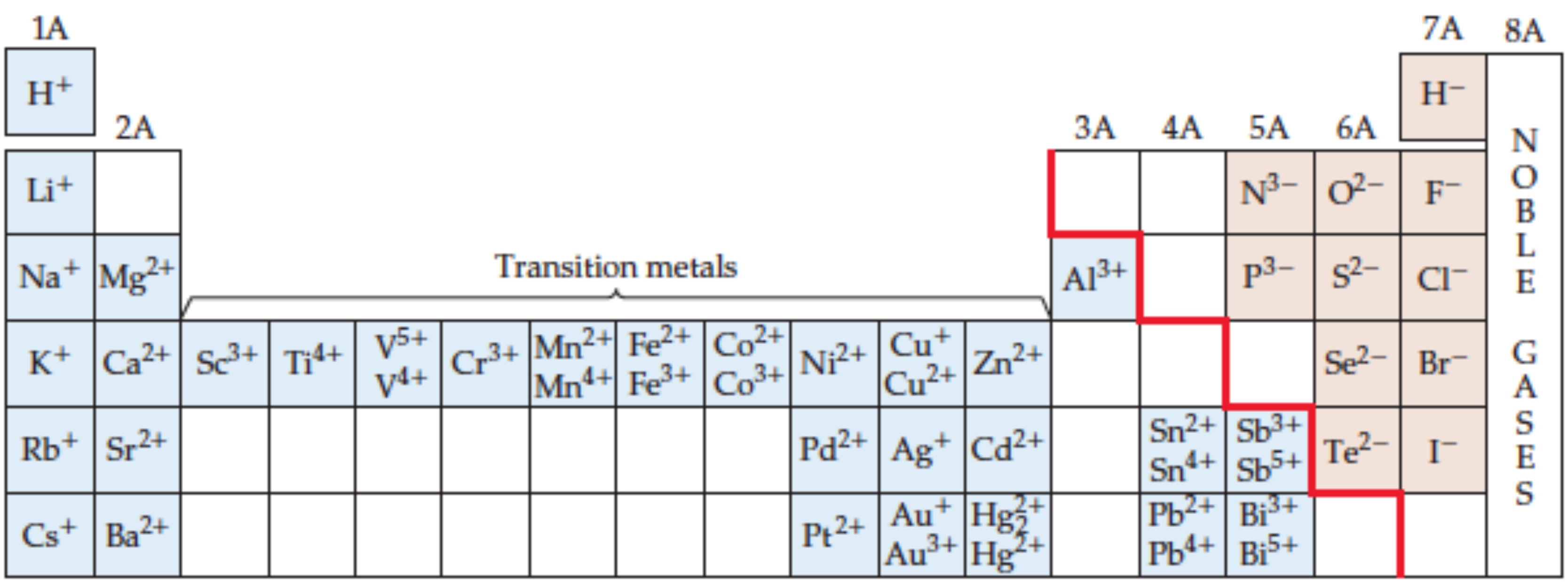
F : 9 P^+ can pull its 9 e^- closer.

F is smaller than F^-



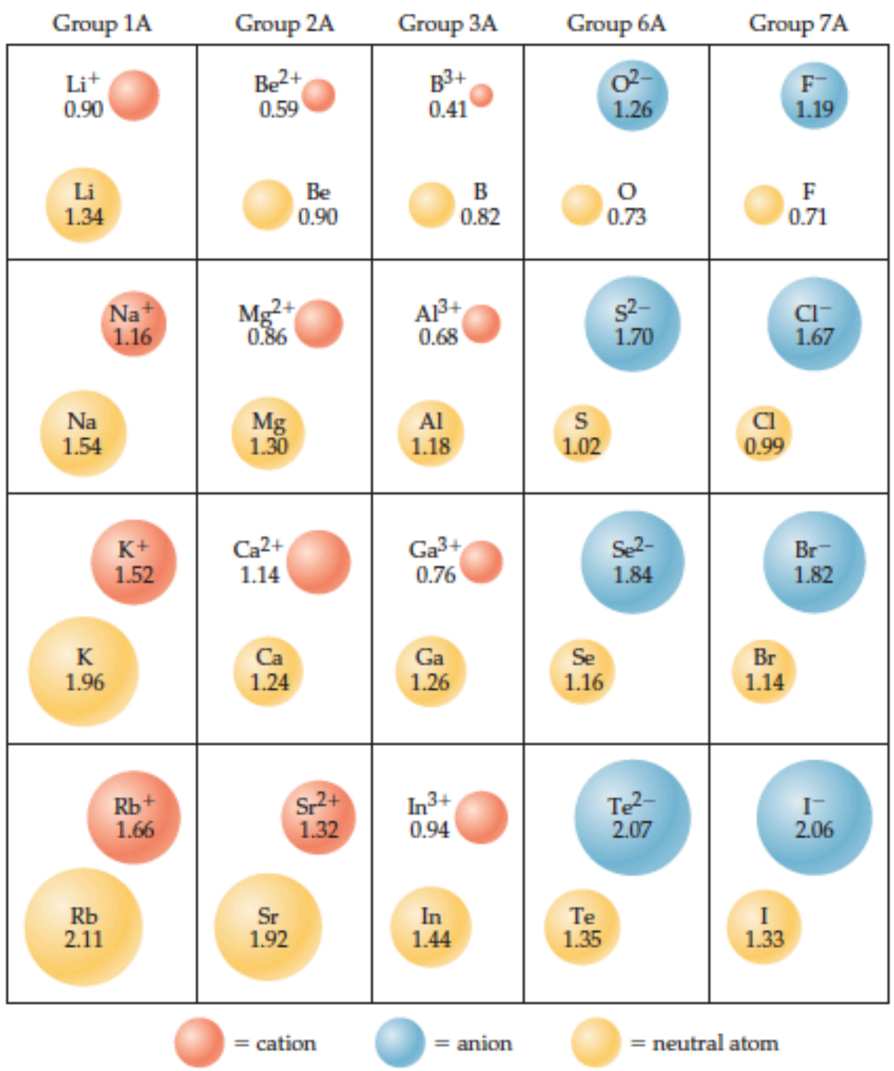
fluoride radii = 1.19 angstroms

In the table (below), there are examples of some stable ions.
The ions of groups 1A (X^{+}), 2A (X^{2+}), 3A (X^{3+}), 5A (X^{3-}), 6A (X^{2-}), 7A (X^{-}), have an electron configuration like that of a noble gas, usually the closest to it.



Some stable ions do not have noble gas configurations. The transition metals often form ions without complete octet. These ions are all cations. Sometimes, one element can form cations with more than one charge such as iron (Fe^{2+}, Fe^{3+}).

The table below shows that the cations (orange color) are smaller than their parent atoms. likewise, the anions (blue color) are larger than that their parent atoms. It also shows how the radius of neutral atoms (yellow color) increases as you down a group (column) and decreases as you move across a period (row).



Summary

Group #	Ion Formed	Equation	Example (name)
Group 1A	X^+	$X - 1\ e^- = X^+$	**K** $^+$ (potassium ion)
Group 2A	X^{2+}	$X - 2\ e^- = X^{2+}$	**Mg** $^{2+}$ (magnesium ion)
Group 3A	X^{3+}	$X - 3\ e^- = X^{3+}$	**Al** $^{3+}$ (Aluminum ion)
Group 5A	X^{3-}	$X + 3\ e^- = X^{3-}$	**N** $^{3-}$ (nitride)
Group 6A	X^{2-}	$X + 2\ e^- = X^{2-}$	**O** $^{2-}$ (oxide)
Group 7A	X^-	$X + 1\ e^- = X^-$	**F** $^-$ (fluoride)