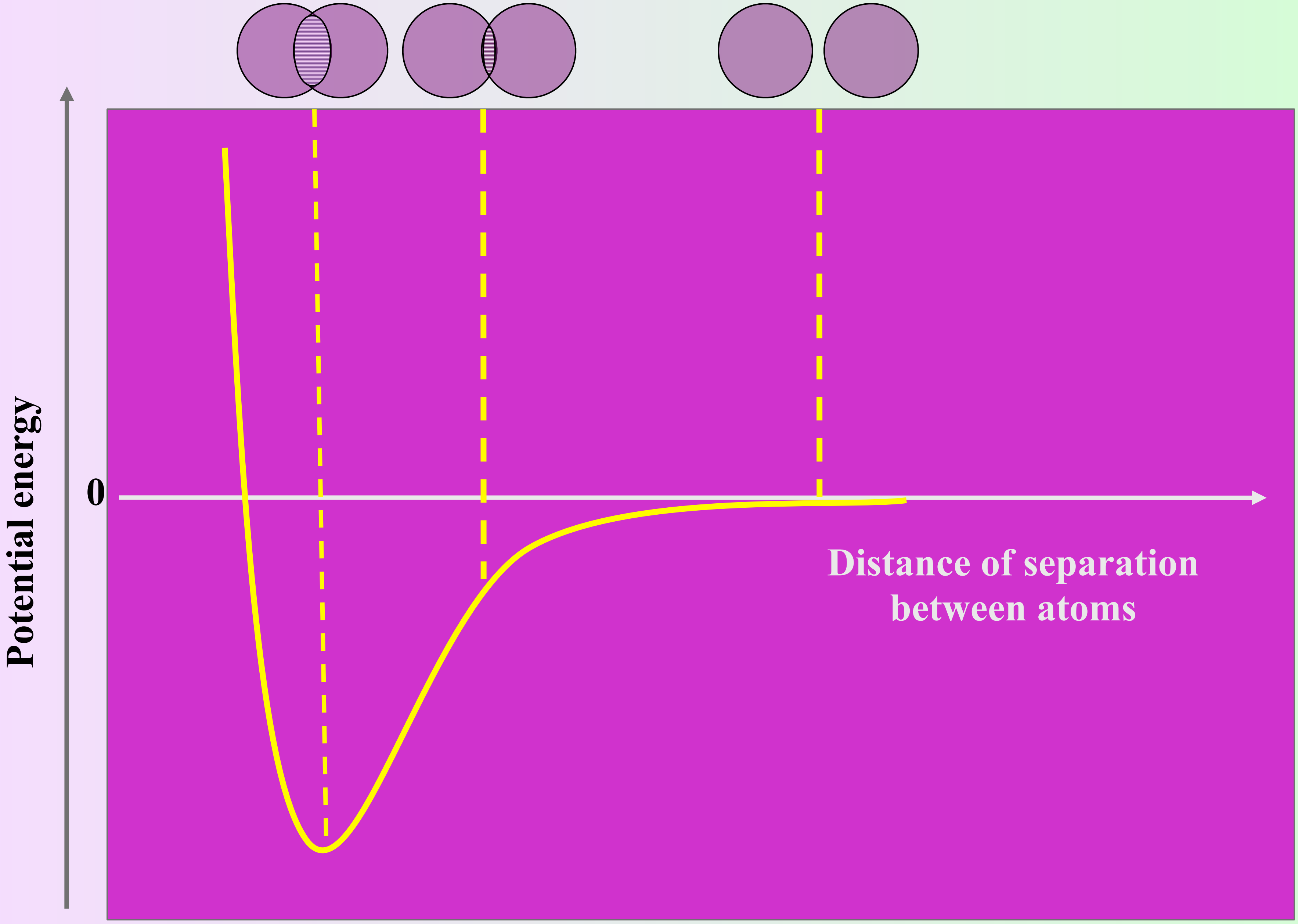
Forming Chemical Bonds

Why do atoms form chemical bonds ?

so that the system can achieve the lowest possible potential energy

Example covalent bonding in H_2

H •   • H



Potential energy
0
Distance of separation
between atoms

Ions and Ionic Compounds

Ionic compounds

Any compound when melted that conducts electricity is considered ionic

Example : $NaCl_{(s)} \longrightarrow NaCl_{(l)}$

$NaCl_{(s)}$ is made up of Na^+ and Cl^- ions

Ionic Bonds

especially prevalent in compounds formed between group 1A and 2A elements with group 6A and 7A elements.

between Elements with the biggest difference in electronegativity

Ionic and Molecular Compounds

Ionic compounds are usually formed between metals and nonmetals.

Molecular compounds are usually formed between two nonmetals.

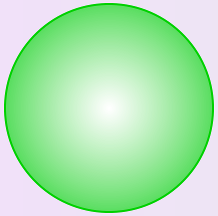
Ions

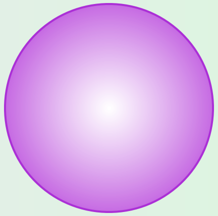
When electrons are removed from or added to a neutral atom or molecule, a charged particle called an ion is formed.

Positively charged ions are called cations

Negatively charged ions are called anions

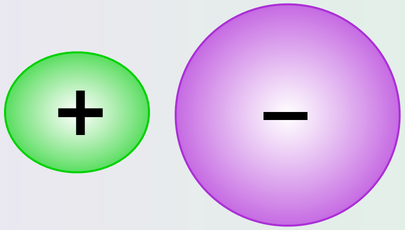
Atoms vs Ions





Na: 11 protons;
11 electrons

Cl: 17 protons;
17 electrons



Na^+: 11 protons;
10 electrons

Cl^-: 17 protons;
18 electrons

Oxidation number

also called ***oxidation state***

signifies the number of charges the atom would have in a molecule (or an ionic compound) if electrons were transferred completely.

Octet rule

Elements (*other then hydrogen*) tend to react to acquire a noble gas electron configuration.

Electron Configurations of Cations and Anions

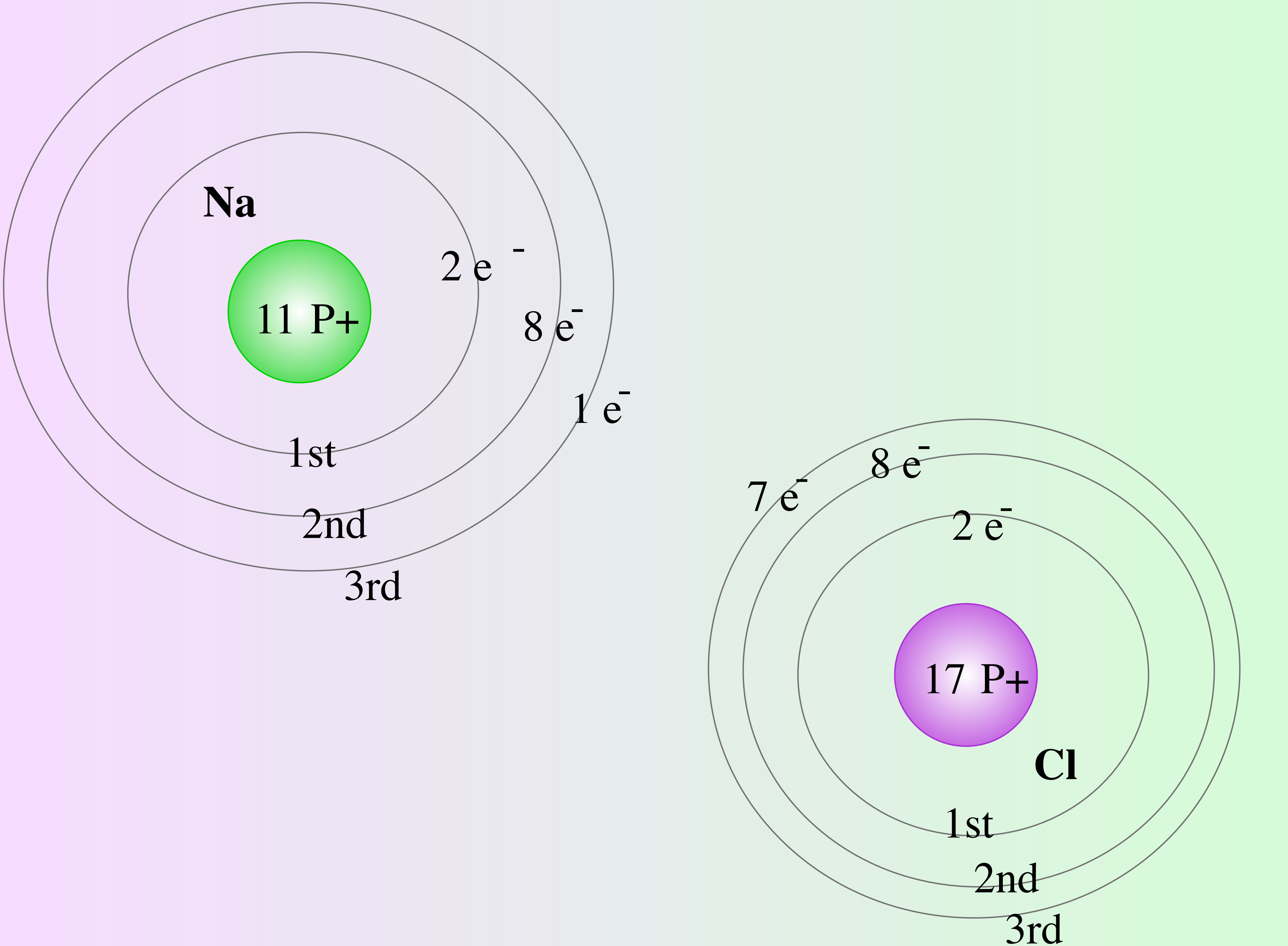
Element	Atom		Ion	
Sodium	Na	[Ne]$3s^1$	Na^+	[Ne]
Magnesium	Mg	[Ne]$3s^2$	Mg^{2+}	[Ne]
Aluminum	Al	[Ne]$3s^23p^1$	Al^{3+}	[Ne]

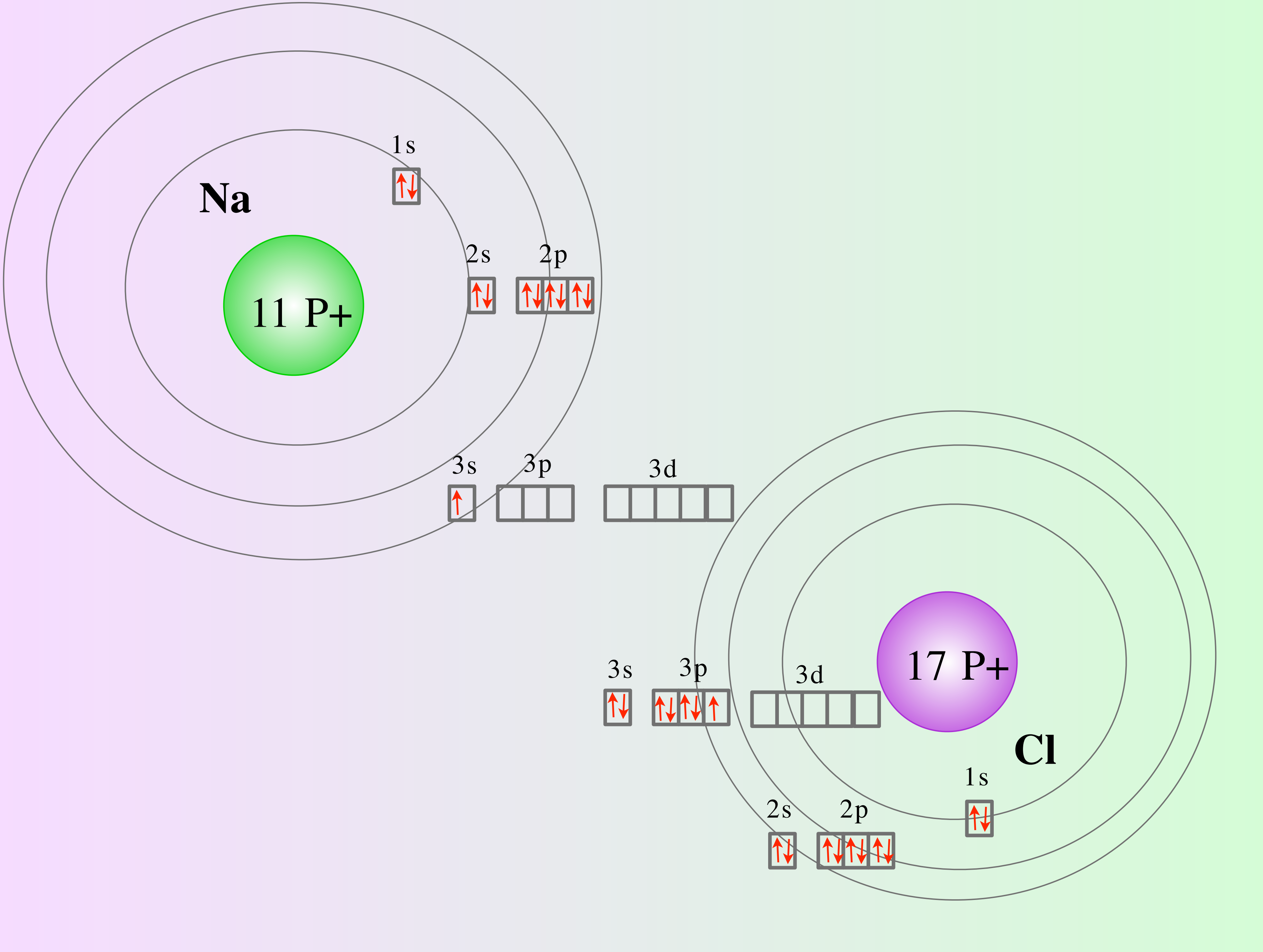
these cations now have a complete octet

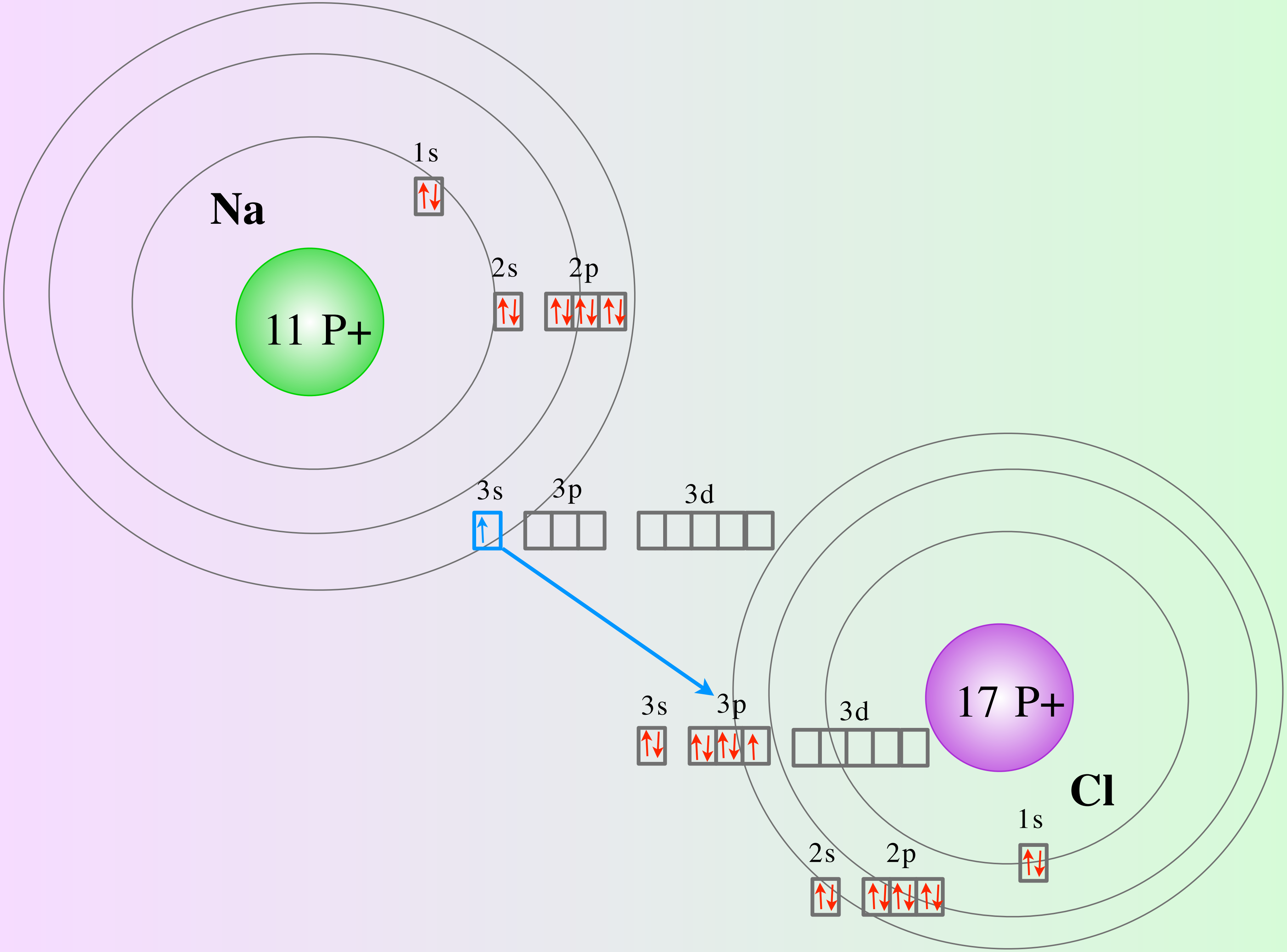
Electron Configurations of Cations and Anions

Element	Atom		Ion	
Hydrogen	H	$1s^1$	H^-	$1s^2$
Fluorine	F	$1s^2 2s^2 2p^5$	F^-	$1s^2 2s^2 2p^6$
Oxygen	O	$1s^2 2s^2 2p^4$	O^{2-}	$1s^2 2s^2 2p^6$
Nitrogen	N	$1s^2 2s^2 2p^3$	N^{3-}	$1s^2 2s^2 2p^6$

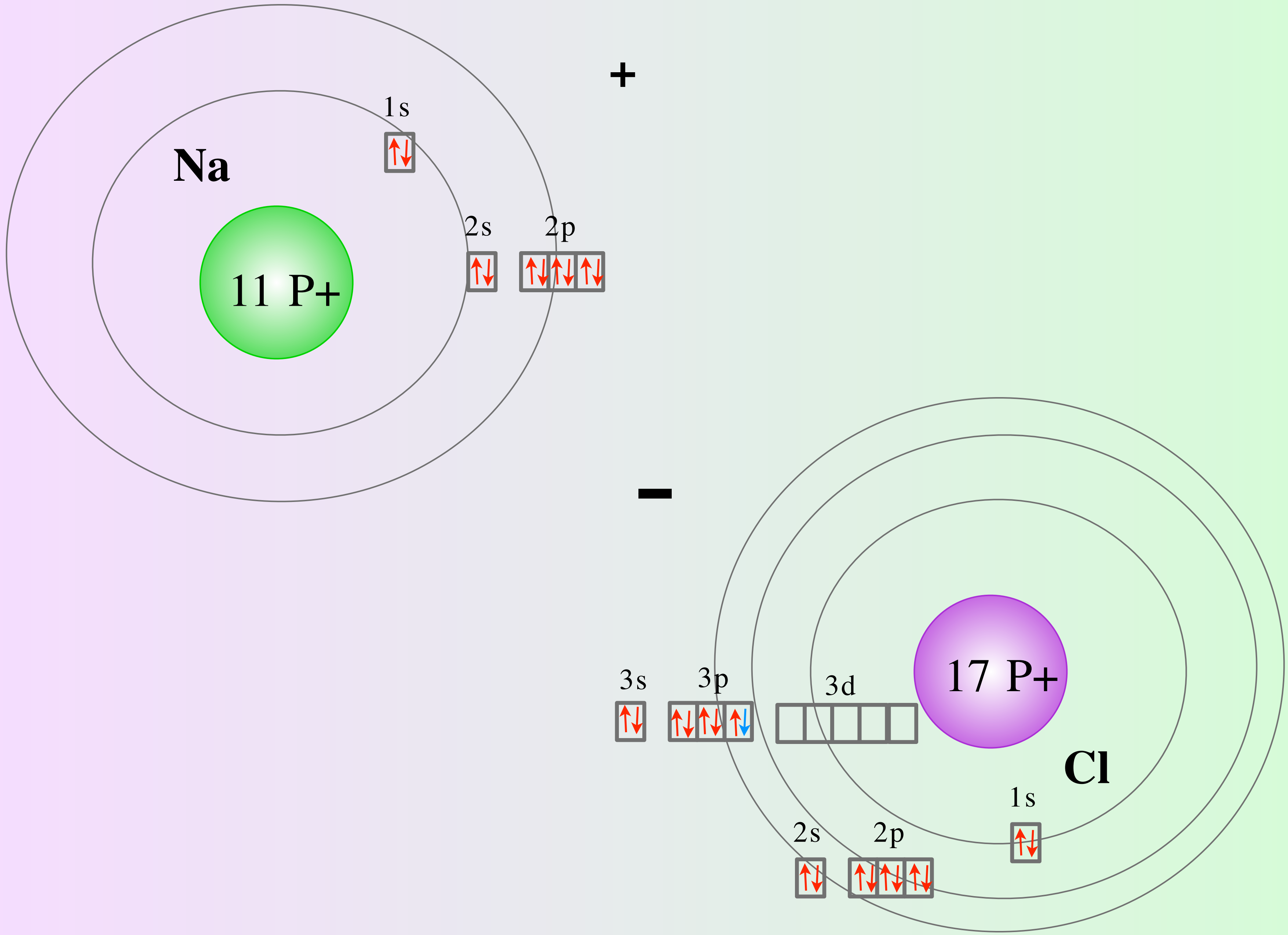
these anions now have a complete octet


Na
11 P+
2 e-
8 e-
1 e-
1st
2nd
3rd
7 e-
8 e-
2 e-
17 P+
Cl
1st
2nd
3rd

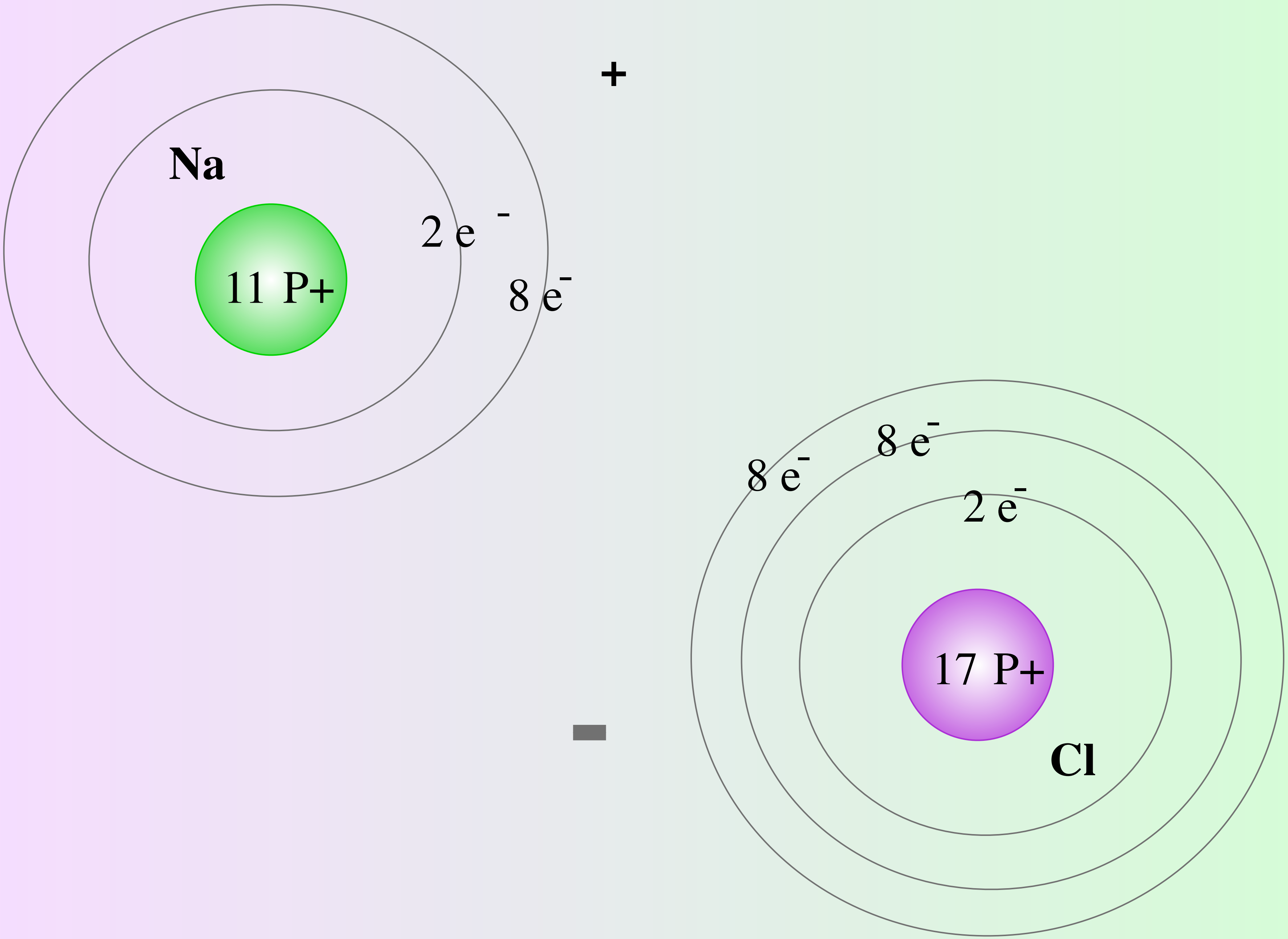

Na
11 P+
1s
2s
2p
3s
3p
3d
17 P+
Cl
3s
3p
3d
1s
2s
2p



Na
11 P+
1s
2s
2p
3s
3p
3d
17 P+
Cl
3s
3p
3d
1s
2s
2p



+
Na
11 P+
1s
2s
2p
–
3s
3p
3d
17 P+
Cl
1s
2s
2p



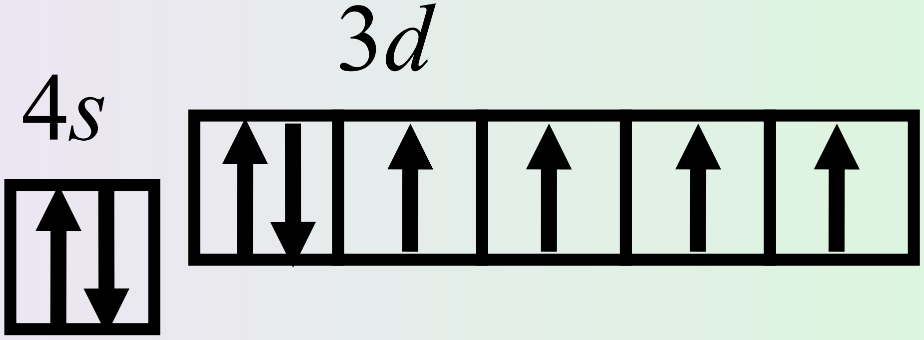
+
Na
11 P+
2 e -
8 e -
-
8 e -
8 e -
2 e -
17 P+
Cl

Transition Metals

commonly lose their two valence electrons and possibly some electrons from their *d* subshell

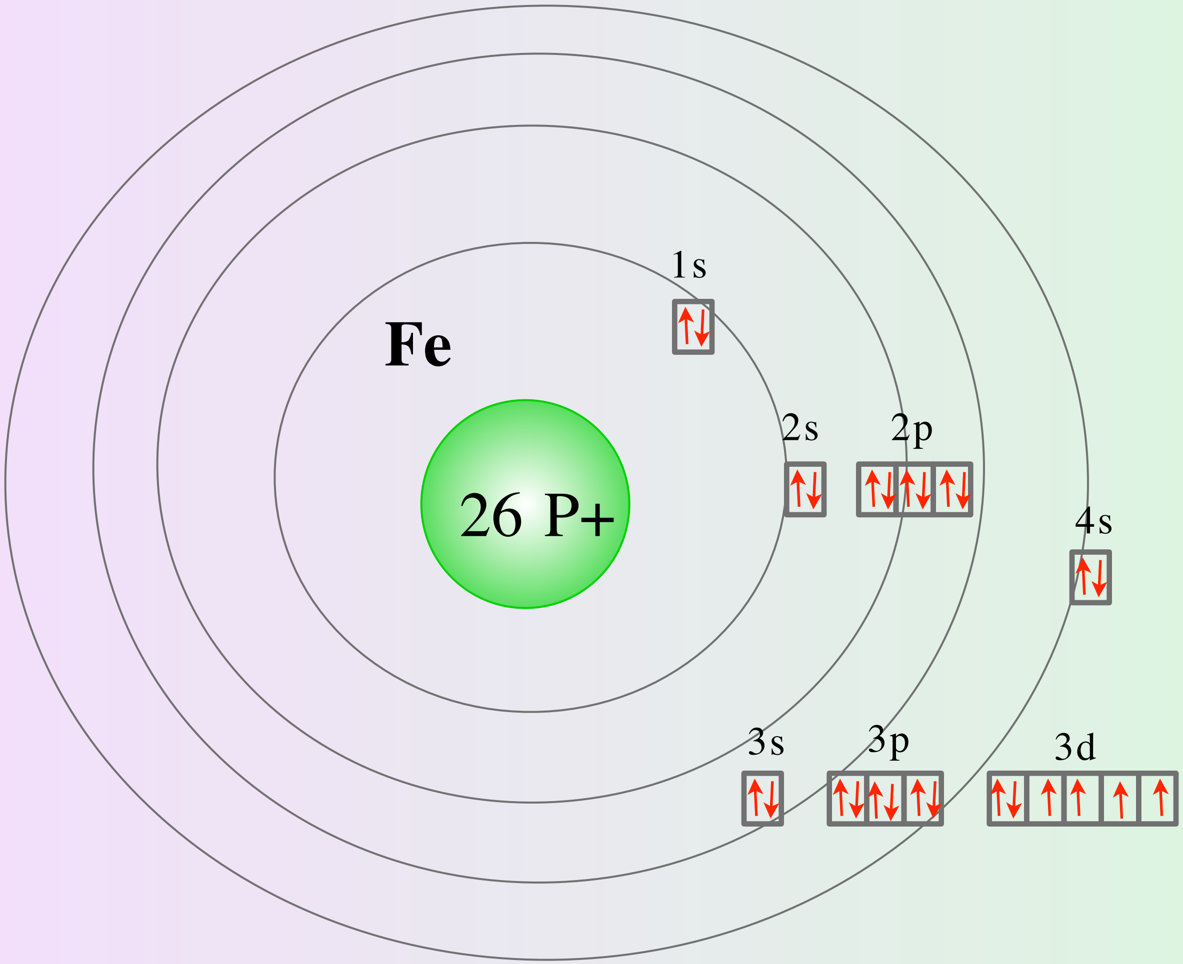
$_{26}$ Fe

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

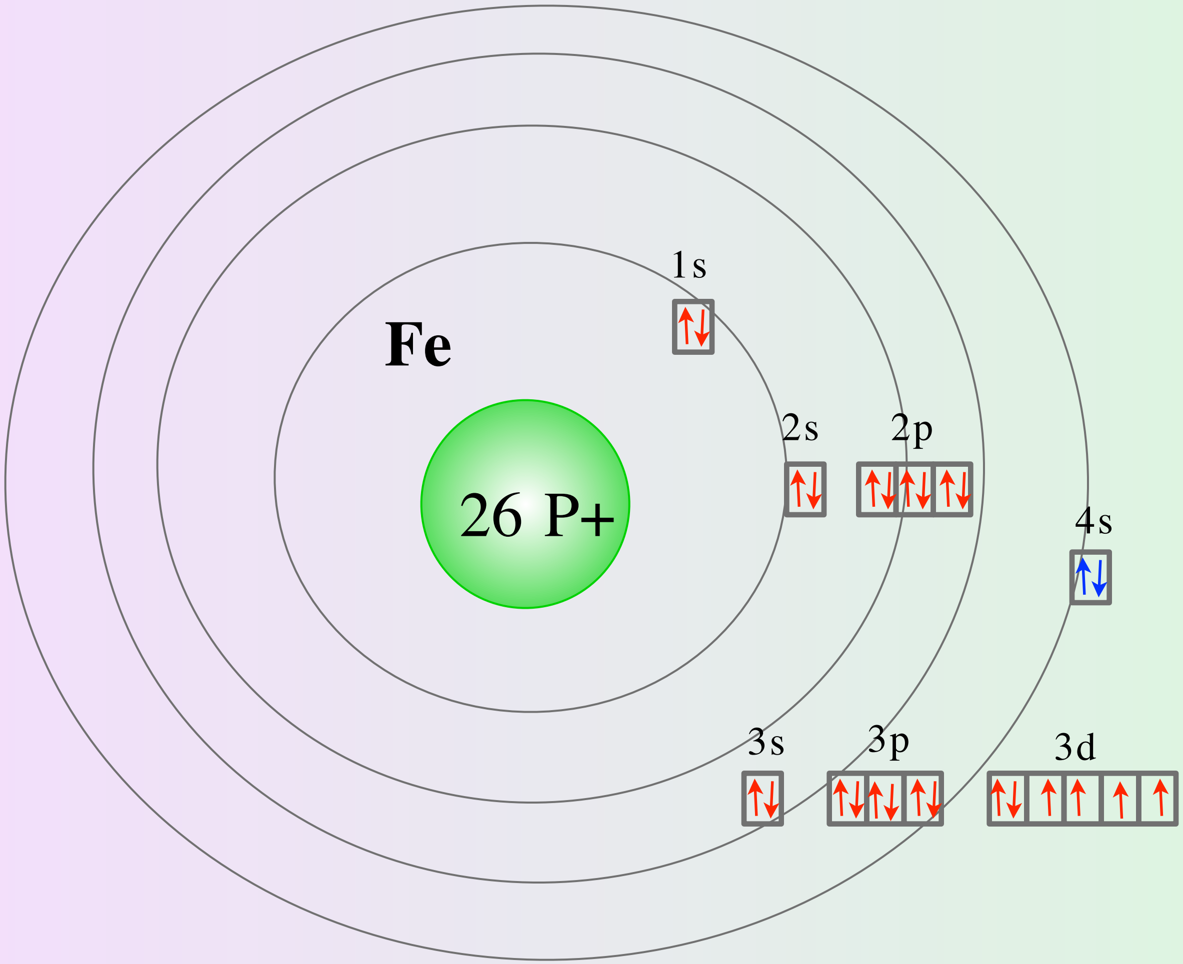


[Ar]

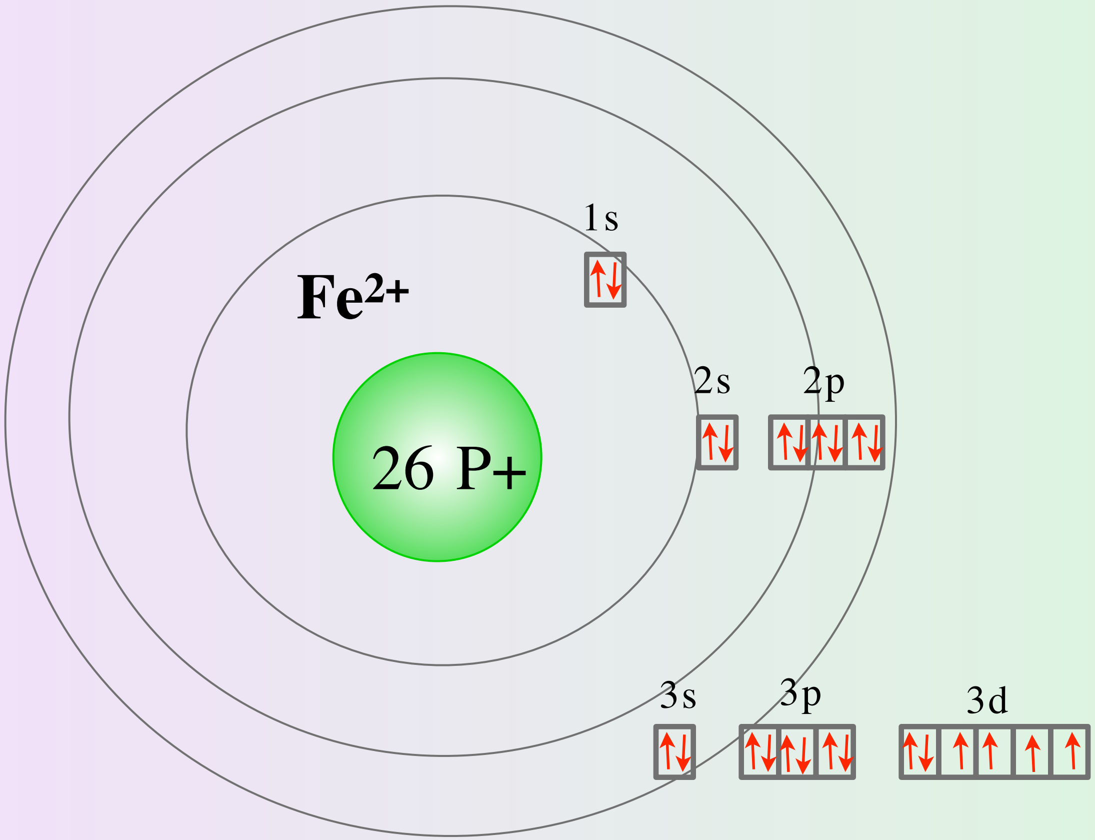
$_{26}$ Fe



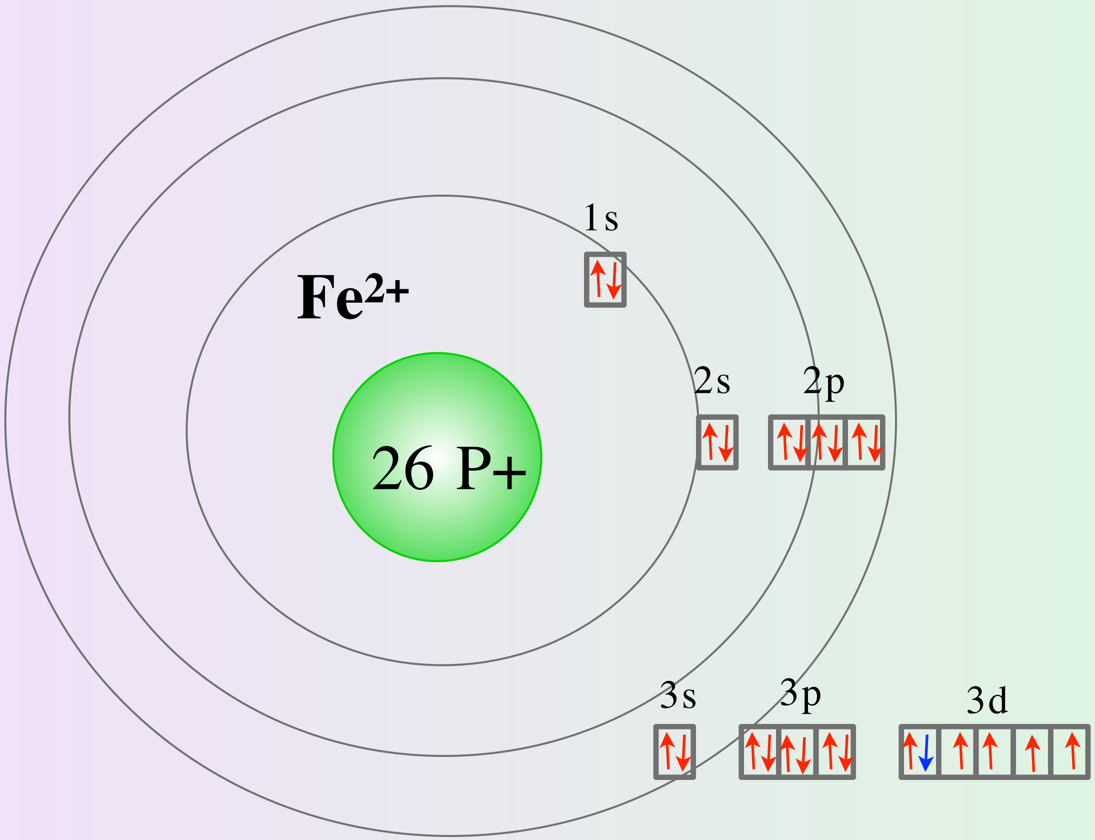
$_{26}$ Fe



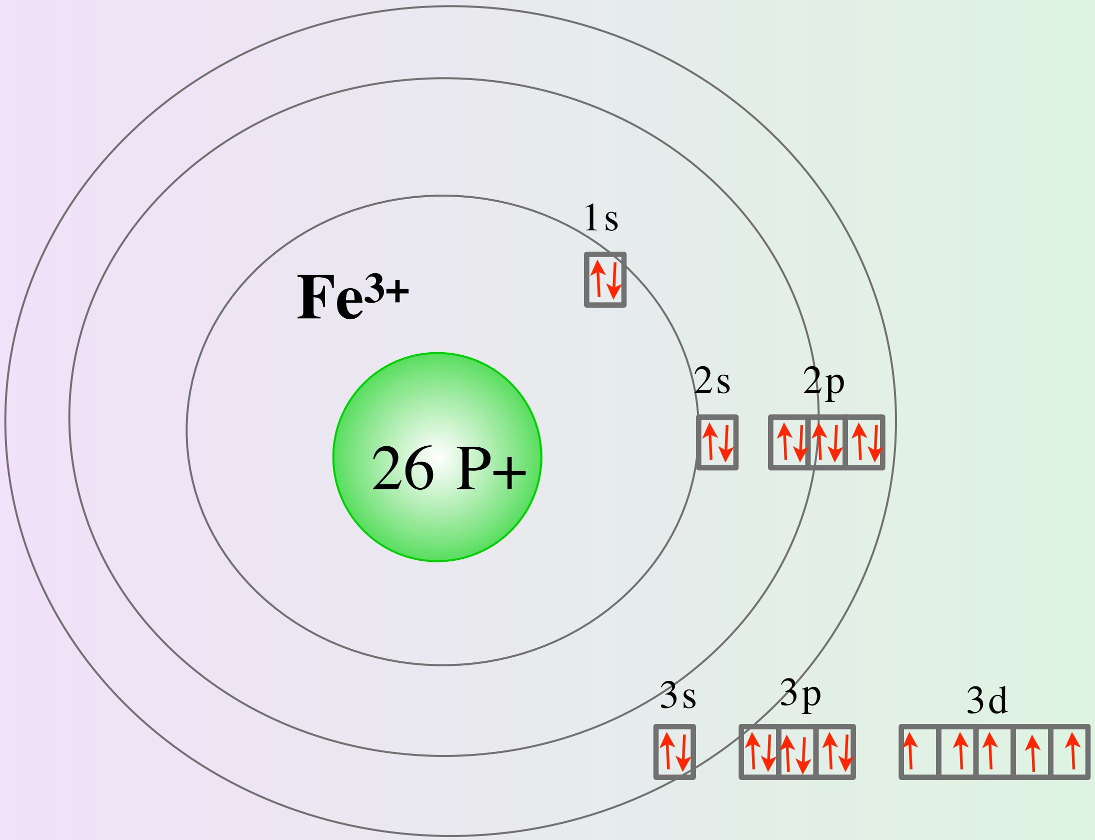
$_{26}Fe^{2+}$



$_{26}\,Fe^{2+}$



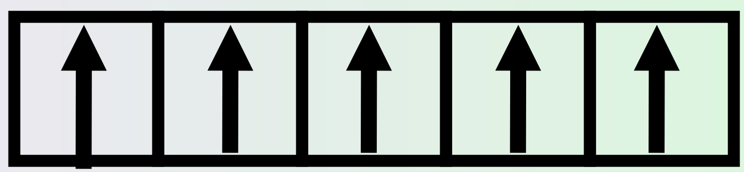
$_{26}\,Fe^{3+}$



$_{26}Fe^{3+}$

[Ar] 3d^5

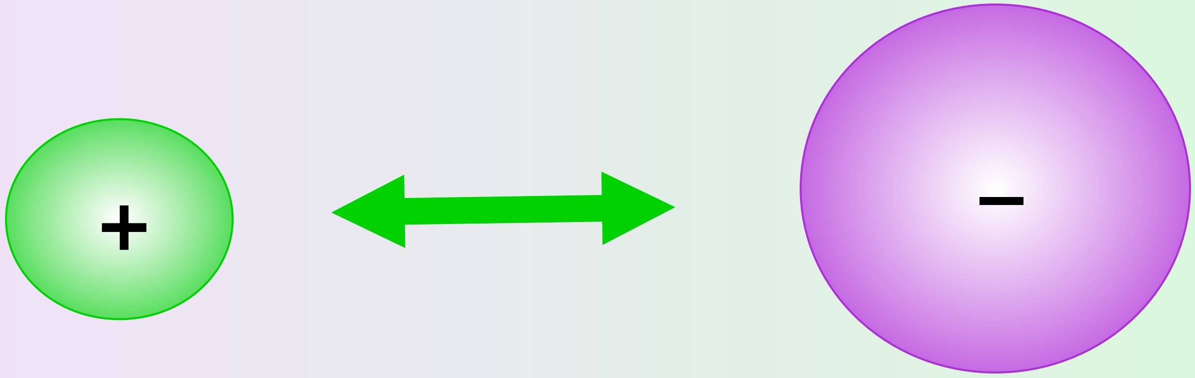
3*d*



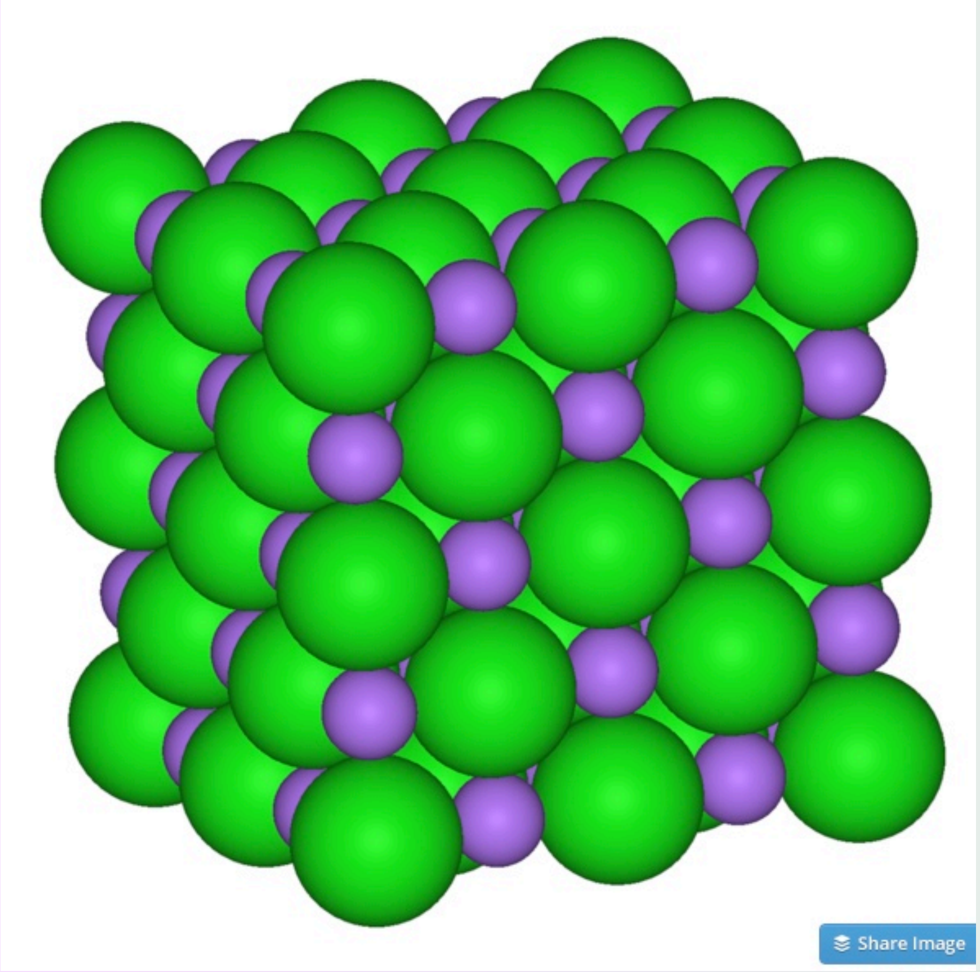
[Ar]

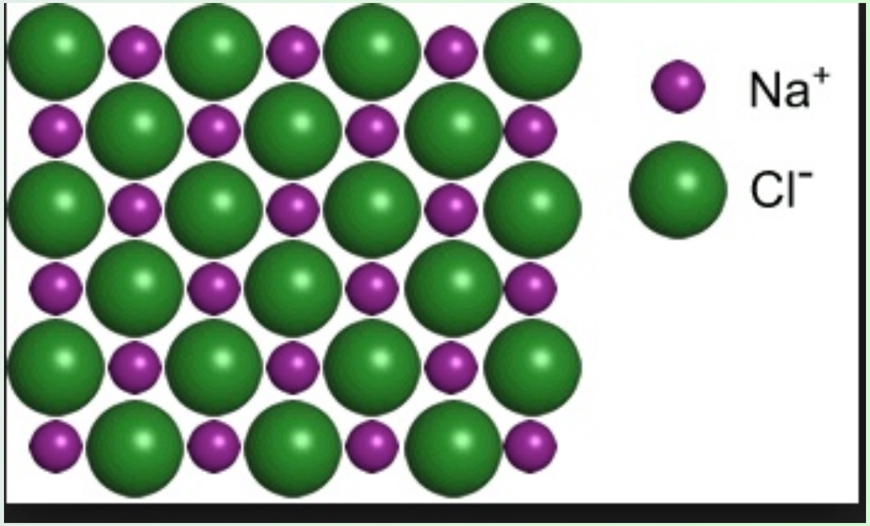
Ionic Bonding

electrostatic attraction between oppositely charged ions



Structure of solid NaCl.





Ionic Compounds

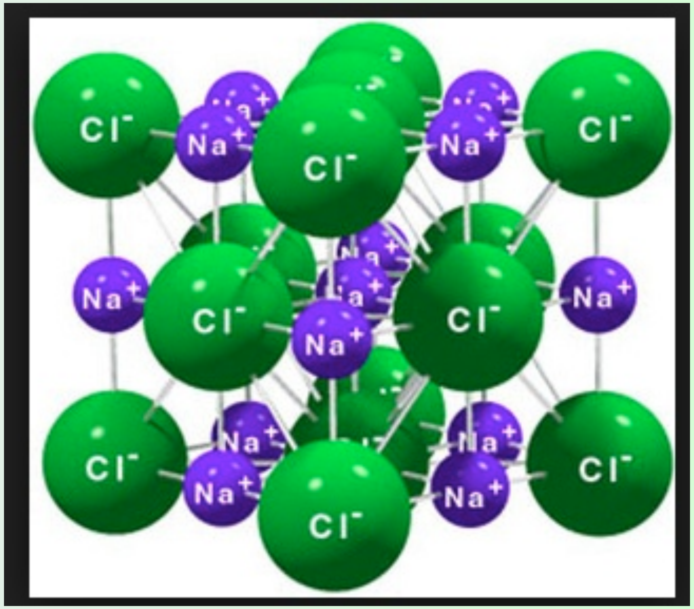
discrete molecules are not present, so ionic compounds are represented by their **empirical formulas**

some times referred to as **formula units**

Empirical Formula

The empirical formula tells us which elements are present and the simplest whole-number ratio of their atoms.

Ionic Crystals



high melting points

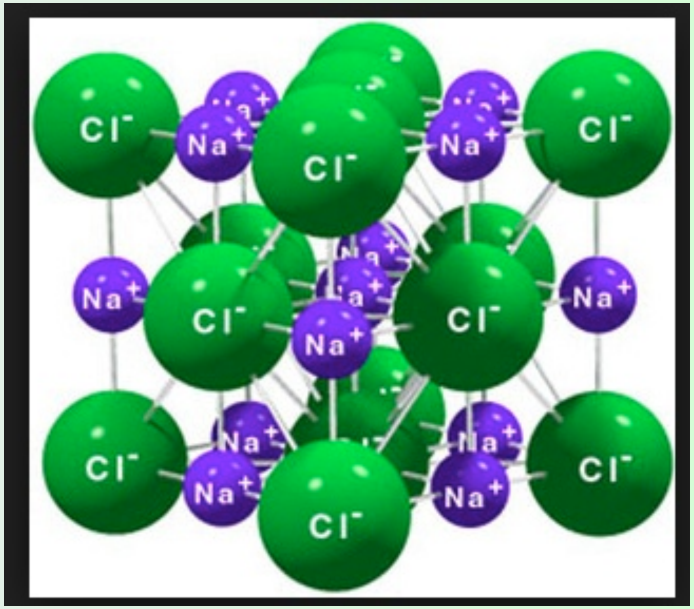
high boiling points

hard, rigid and brittle solids

NaCl

melting point 801 ^{o}C

boiling point 1413 ^{o}C

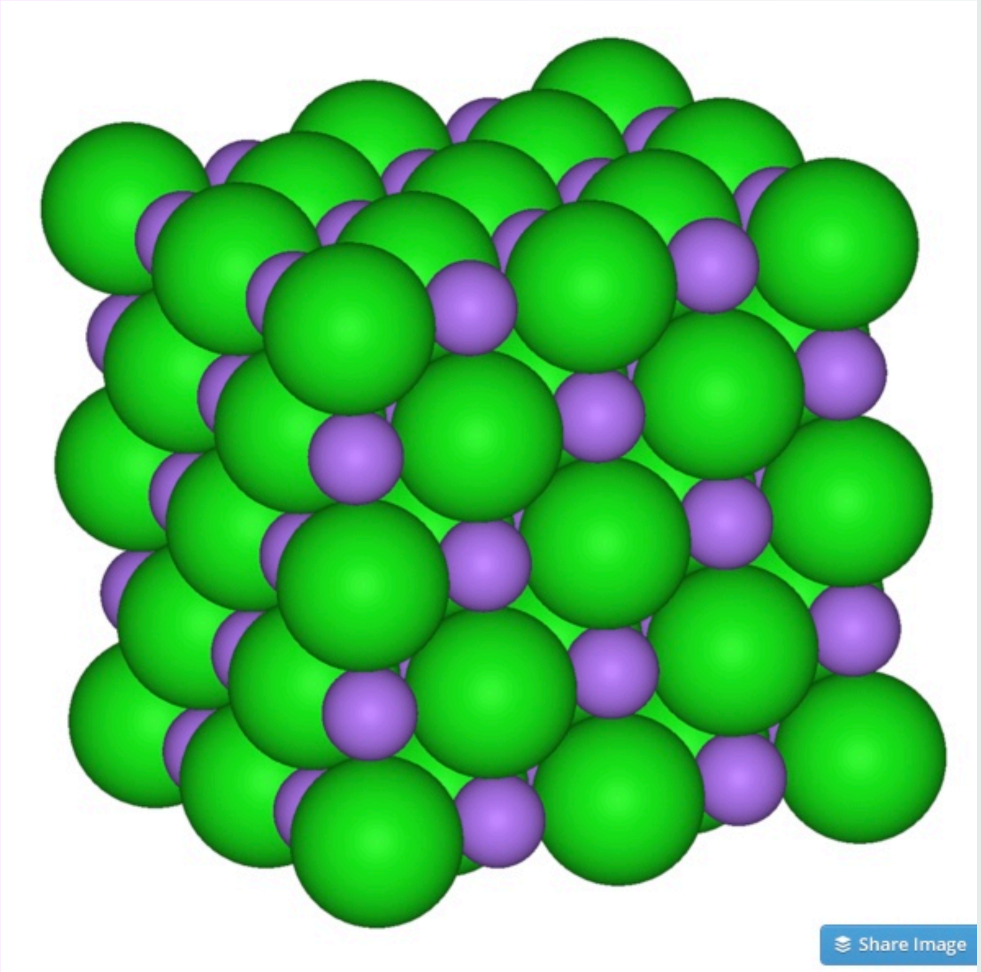


Formation of LiF

$$Li(s) + 1/2\ F_2(g) \longrightarrow LiF(s)$$

$$Li\cdot + \cdot\ddot{\underset{..}{F}}: \longrightarrow Li^+ \quad [:\ddot{\underset{..}{F}}:]^-$$

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



Solid structure of LiF

Crystal lattice

the ions are packed together to maximize (+) (-) attractions and minimize (+) (+) and (-) (-) repulsions

Coulomb's Law

Energy of electrostatic attraction is directly proportional to product of charge and inversely proportional to distance

$$\text{lattice energy} = k \frac{Q_1 Q_2}{r}$$

$k = (2.31 \times 10^{-19}$ J nm)

therefore, strong lattices are favored when the ions have a high charge to size ratio

Lattice Energy of Ionic Compounds

Review : Electron affinity

is the energy change that occurs when an electron is accepted by an atom in the gaseous state (kJ/mol).

$$X(g) + e^{-} \longrightarrow X^{-}(g)$$

The more negative the electron affinity ,the greater the tendency of the atom to accept an electron .

Note:

chlorine **has the greatest electron affinity of any element:**

$$Cl(g) + e^- \longrightarrow Cl^-(g) \qquad \Delta H = -349 \text{ kJ}$$

No metal has an ionization energy low enough to make electron transfer to chlorine exothermic

$$Cs(g) \longrightarrow Cs^+(g) + e^- \qquad \Delta H = +382 \text{ kJ}$$

Lattice energy

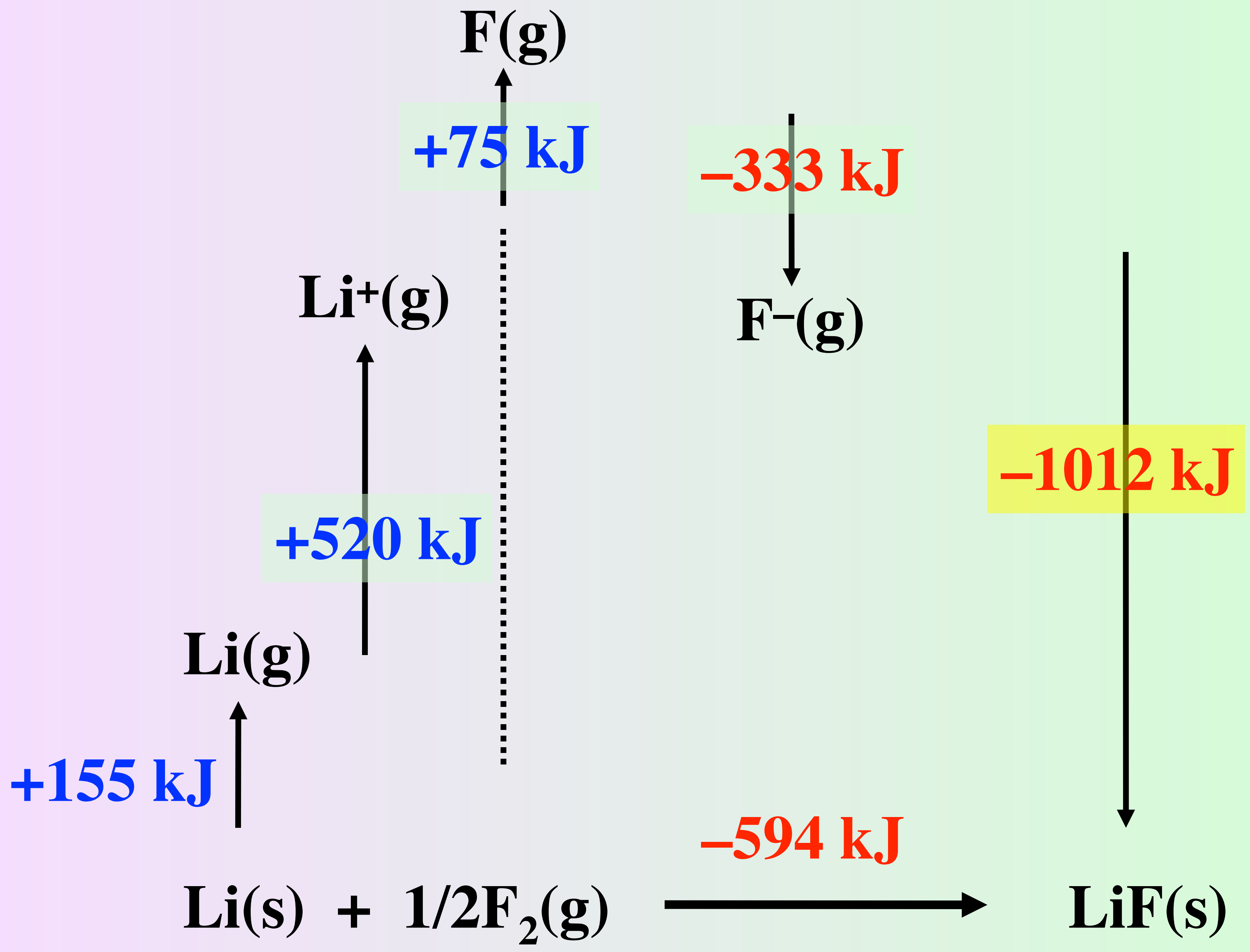
the energy required to completely separate one mole of a solid ionic compound into gaseous ions.

or the energy released when an ionic solid forms from its ions

negative sign (-)

The Born-Haber Cycle for Determining Lattice Energies

Relates lattice energies of ionic compounds to ionization energies, electron affinities, and other atomic and molecular properties.


F(g)
+75 kJ
–333 kJ
Li+(g)
F−(g)
–1012 kJ
+520 kJ
Li(g)
+155 kJ
–594 kJ
Li(s) + 1/2F2(g)
LiF(s)

Formation of ionic compounds

requires the lattice energy to be sufficiently large to overcome ionization energy of the element that forms the cation.

balance between energy input (ionization energies) and stability gained from formation of the solid

Formation of ionic compounds

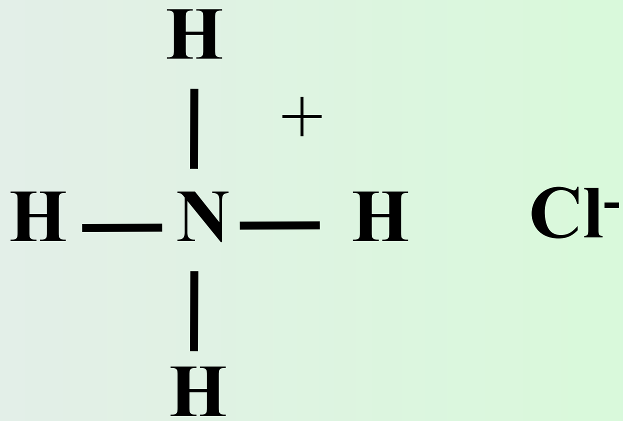
The main impetus for the formation of an ionic compound rather than a covalent compound results from the strong mutual attraction among the ions

Polyatomic ions

Ionic bonding involving polyatomic ions is some what ambiguous

example (ammonium-chloride) NH_4Cl

being that NH_4^+ is held together by covalent bonds



Example

bonding in $MgCl_2$ is ionic;

bonds in $AlCl_3$ are polar covalent

Ionization energies (I.E.)

$$Mg(g) \longrightarrow Mg^{2+}(g) + 2e^-$$

$$\Delta H = I.E._1 + I.E._2$$

$$= 753 + 1435 = +2180 \text{ kJ/mol}$$

$$Al(g) \longrightarrow Al^{3+}(g) + 3e^-$$

$$\Delta H = I.E._1 + I.E._2 + I.E._3$$

$$= 580 + 1815 + 2740 = +4125 \text{ kJ/mol}$$

Example

bonds in $AlCl_3$ are polar covalent

energy input (ionization energies) out weighs stability gained from formation of an ionic solid

$$Al\,(g) \longrightarrow Al^{3+}(g) + 3e^-$$

$$\Delta H = I.E._1 + I.E._2 + I.E._3$$

$$= 580 + 1815 + 2740 = +4125 \text{ kJ/mol}$$