Matter and material: Chemical bonding

Objectives

- Define a chemical bond.
- Draw Lewis dot diagrams of elements.
- Define a covalent bond.
- Draw Lewis dot diagrams of covalent molecules.
- Define an ionic bond.
- Draw Lewis dot diagrams of half reactions, full reactions and final products of ionic compounds.
- Learn the common cations and anions not identified on the periodic table.
- Define a metallic bond.
- Calculate the relative atomic mass for covalent molecules and the relative formula mass for ionic compounds.

Introduction to chemical bonds

A chemical bond is the physical process that causes atoms to be attracted to each other and held together. Atoms are held together in a bond by an electrostatic force of attraction between positively charged protons and negatively charged electrons.

A chemical bond [D] is a strong electrostatic force of attraction that holds atoms or ions together, resulting from the simultaneous attraction between their nuclei and electrons.

Depending on the atoms involved, there will either be a covalent bond, an ionic bond or a metallic bond.

- A **covalent bond** exists between nonmetals. Note that a metalloid and a nonmetal also bond covalently. The atoms share electrons. This results in an electrostatic force of attraction since the shared electrons are attracted to both nuclei.
- An **ionic bond** exists between a metal and a nonmetal. Electrons are transferred from the metal to the nonmental. The metal becomes positively charged while the nonmetal becomes negatively charged resulting in an electrostatic force of attraction between the atoms.
- A **metallic bond** exists in a metal or a mixture of metals. The valence electrons of the metal atoms become delocalised and form a sea of electrons. This results in an electrostatic force of attraction between the delocalised electrons and the positive metal ion core.

Valence electrons and electron dot diagrams

We already learnt that atoms have core electrons and valence electrons. The valence electrons determine the properties of an element since those are the electrons involved in chemical reactions. Based on the arrangement of the elements in the periodic table, the number of valence electrons corresponds with the group number:

- Group 1 elements have 1 valence electron.
- Group 2 elements have 2 valence electrons.
- Group 13 elements have 3 valence electrons.
- Group 14 elements have 4 valence electrons.
- Group 15 elements have 5 valence electrons.
- Group 16 elements have 6 valence electrons.
- Group 17 elements have 7 valence electrons.
- Group 18 elements have 8 valence electrons, except for helium which has 2 valence electrons.

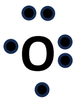
Lewis notation is a system of dots and crosses to represent valence electrons of an atom. The symbol for the element is written in the middle, and dots are drawn around the symbol. The first four dots are drawn as one on each side, usually starting at the top of the symbol and moving in a clockwise direction. If there are more than four valence electrons, they pair up with the first four.

For example, hydrogen is in group 1 so it has 1 valence electron. This is the Lewis dot diagram of hydrogen:

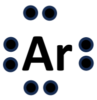
Magnesium is in group 2 so it has 2 valence electrons. This is the Lewis dot diagram of magnesium:

Helium also has 2 valence electrons, but because it has a full outer energy level, we represent the Lewis dot diagram of helium in a way where we pair up the 2 valence electrons:

Oxygen is in group 16 so it has 6 valence electrons. This is the Lewis dot diagram of oxygen:



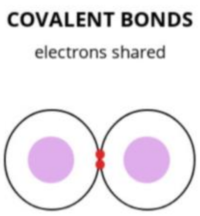
Argon is in group 18 and has 8 valence electrons. This is the Lewis dot diagram of argon:

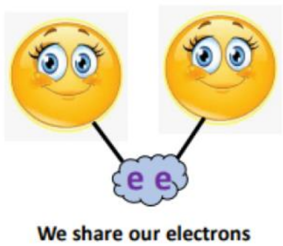


Most elements combine to form molecules so that they have 8 valence electrons. Having the full 8 electrons in its outermost energy level makes the element within the molecule stable. One exception is hydrogen, which combines to form molecules so that it has 2 valence electrons in its outermost energy level.

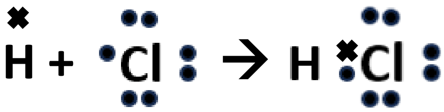
<u>Covalent bonding</u>

A covalent bond [D] is the sharing of electrons between nonmetals.

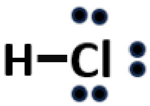




Let's consider the covalent bond in hydrogen chloride. Hydrogen has 1 valence electron. The first energy level can only fit two electrons, so hydrogen only needs one more electron to have a full outer energy level. All other elements need 8 valence electrons to have a full outer energy level. Chlorine has 7 valence electrons, so it needs one more electron to have a full outer energy level. We use a Lewis dot diagram to help us represent the bond between the atoms:

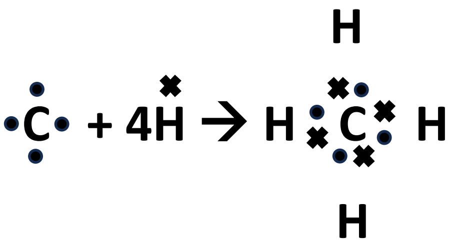


As we can see, hydrogen is sharing one electron and chlorine is sharing one electron. These two shared electrons make up a <u>single bond</u> which may also be represented as follows:

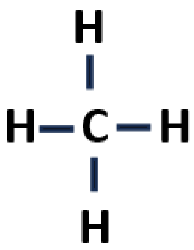


The single line represents that each atom involved in the bond shares one of its electrons. Note that the total number of valence electrons does not change. The electrons are shared so that both H and Cl have full outer energy levels.

Let's consider the covalent bond in carbon tetrahydride, CH_4, commonly known as methane. Hydrogen has 1 valence electron, so it needs one more electron to have a full outer energy level. Carbon has 4 valence electrons, so it needs four more electrons to have a full outer energy level. We use a Lewis dot diagram to help us represent the bonds between the atoms:



Now we can understand why a carbon atom bonds with 4 hydrogen atoms. In this case, there are four <u>single bonds</u> around the C atom. A single bond is made up of one shared pair of electrons, in this case one electron from C and one electron from H. This may be represented as follows:



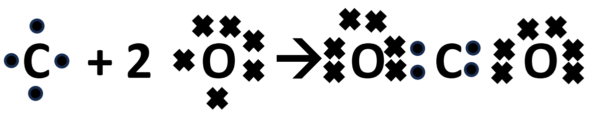
Let's consider the covalent bond in carbon dioxide, CO_2. Carbon has 4 valence electrons, so it needs four more electrons to have a full outer energy level. Oxygen has 6 valence electrons, so it needs two more electrons to have a full outer energy level.

How do we know which atom goes in the centre? <u>Hydrogen always goes on the ends</u>. Except for hydrogen, the <u>central atom is the one with the lowest electronegativity</u>.

Electronegativity [D] is the tendency of an atom in a molecule to attract bonding electrons closer to itself.

We will learn more about electronegativity in Grade 11. You may derive this value from the periodic table. It is the number to the left of the element. C has an electronegativity of 2.5 and O has an electronegativity of 3.5. So C will be the central atom.

We now need to think of a way to arrange the electrons around the atoms in such a way that each atom has eight electrons around it – a full outer energy level. We may use single, double or triple bonds to achieve this. Because oxygen needs two more electrons, this suggests a <u>double bond</u>. If an oxygen atom shared two of its electrons with two of carbon's electrons, then it will have eight electrons around it. We use a Lewis dot diagram to help us represent the bonds between the atoms:



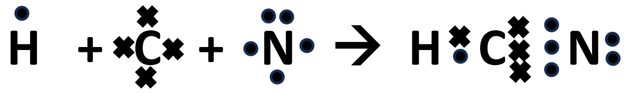
This may also be represented with a double bond on either side of carbon, since two pairs of electrons are shared by each atom:

Note that the rest of the valence electrons that are not involved in bonding need to be **shown** and they always need to be **paired**. One pair of electrons on each oxygen atom should be placed the **furthest** away from the shared electrons due to repulsion. The other pair may be placed above or below each O atom.

Let's consider the covalent bond in HCN. Hydrogen always goes at the end of a molecule. C has an electronegativity of 2.5 and N has an electronegativity of 3. So C will be the central atom.

Hydrogen has 1 valence electron, so it needs one more electron to have a full outer energy level. Carbon has 4 valence electrons, so it needs four more electrons to have a full outer energy level. Nitrogen has 5 valence electrons, so it needs three more electrons to have a full outer energy level.

We now need to think of a way to arrange the electrons around the atoms in such a way that each atom has eight electrons around it – a full outer energy level. We may use single, double or triple bonds to achieve this. Because hydrogen needs one more electron, this suggests a <u>single bond</u>. Because nitrogen needs three more electrons, this suggests a <u>triple bond</u>. We use a Lewis dot diagram to help us represent the bonds between the atoms:



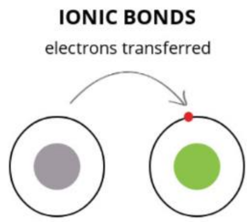
This may also be represented with a single bond between the carbon atom and the hydrogen atom, and a triple bond between the carbon atom and the nitrogen atom.

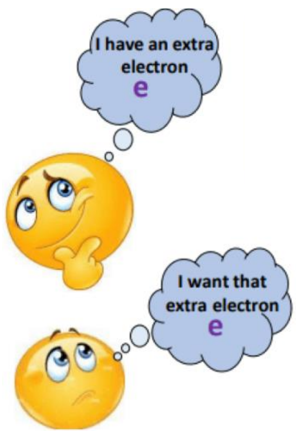
Note that the rest of the valence electrons that are not involved in bonding need to be **shown** and they always need to be **paired**. Nitrogen's two leftover valence electrons should be placed together and the **furthest** away from the shared electrons due to repulsion.

Another important note is that a triple bond is stronger than a double bond, which is stronger than a single bond.

<u>Ionic bonding</u>

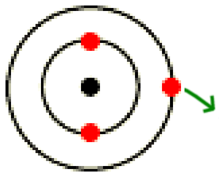
An ionic bond [D] is the transfer of electrons between a metal and a nonmetal.



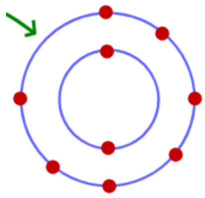


An ion is an element that has lost or gained one or more electrons. It therefore either has a positive charge or a negative charge. Ionic bonds occur when elements transfer electrons so that both elements have a full outer energy level. The overall ionic compound is always neutral.

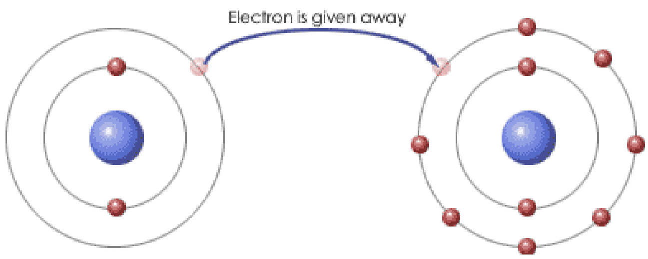
For example, lithium has 1 valence electron. It wants to give up that valence electron to have a full outer energy level:



Fluorine has 7 valence electrons. It therefore wants to gain an electron to have a full outer energy level:



Lithium transfers its valence electron to fluorine. Lithium now becomes Li^+ (lithium cation) and fluorine now becomes Fl^- (fluoride anion). This causes the two atoms to be electrostatically attracted to each other and form LiF (lithium fluoride).



We use Lewis dot diagrams to represent the two half reaction, full reaction and final product of ionic bonds as follows:

Half reaction 1:

×
Li → Li^+ + e-

Half reaction 2:

●● ●● -
F ● + e- → x F ●
● ● ● ●
●● ●●

Full reaction:

× ●● ●● -
Li + F ● → Li^+ + × F ●
● ● ● ●
●● ●●

Final product:

●● -
Li^+ (× F ●)
● ●
●●

Here is another example for potassium sulfite (K_2S):

Half reaction 1:

×
K → K^+ + e-

Half reaction 2:

●●   ●●   2-
● S ● + 2e- → × S ●
● ●   ● ●
●   ×●

In this case, the number of electrons is not balanced. Potassium can give away 1 electron, but sulfur needs 2 electrons. We therefore need to balance out the number of electrons to get to the ratio in which the elements in this compound exist. This means that we need to balance out the first half reaction by multiplying it by 2 to balance out the number of electrons in both equations:

Half reaction 1 balanced:

×
2 K → 2 K^+ + 2e-

Half reaction 2 remains the same:

●●   ●●   2-
● S ● + 2e- → × S ●
● ●   ● ●
●   ×●

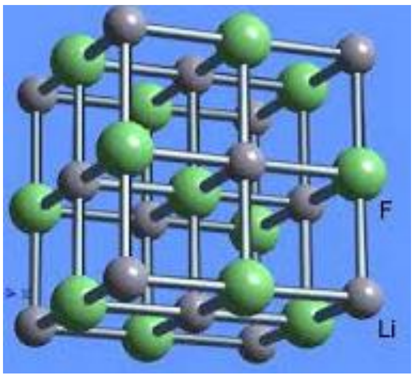
Full reaction:

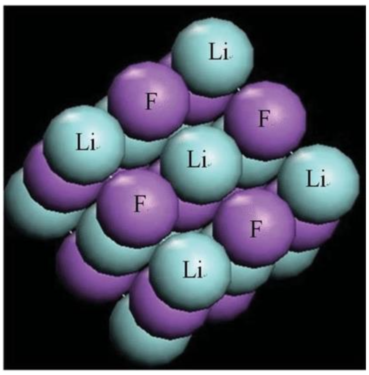
×   ●●   ●●   2-
2 K + S ● → 2 K^+ + × S ●
      ● ●   ● ●
      ●   ×●

Final product:

2K^+ (●● × S ● ● ● ×●) 2-

Ionic compounds exist as solids. The positive and negative ions attract one another so that they are arranged in a three-dimensional network structure. This is known as a crystal lattice. Below is the microscopic representation of the crystal lattice structure of LiF.





Common anions and cations

We can identify the chemical formula of the ions of many elements based on their position on the periodic table. However, there are many common anions and cations. These are either d-block ions or they exist as ionic molecules. The names and formulae, including the charges, need to be memorised as per the table below:

Common cations		Common anions	
Ammonium	NH_4^+	Hydroxide	OH^-
Silver	Ag^+	Hydride	H^-
Copper(I)	Cu^+	Nitrate	NO_3^-
Copper(II)	Cu^{2+}	Nitrite	NO_2^-
Zinc	Zn^{2+}	Sulfate	SO_4^{2-}
Lead(II)	Pb^{2+}	Sulfite	SO_3^{2-}
Nickel	Ni^{2+}	Carbonate	CO_3^{2-}
Tin(II)	Sn^{2+}	Phosphate	PO_4^{3-}
Tin(IV)	Sn^{4+}	Permanganate	MnO_4^-
Mercury(II)	Hg^{2+}	Dichromate	$Cr_2O_7^{2-}$
Iron(II)	Fe^{2+}	Chromate	CrO_4^{2-}
Iron(III)	Fe^{3+}	Hydrogen sulfate	HSO_4^-
Manganese(II)	Mn^{2+}	Hydrogen carbonate	HCO_3^-
Manganese(IV)	Mn^{4+}	Hydrogen sulfite	HSO_3^-
Chromium(III)	Cr^{3+}	Dihydrogen phosphate	$H_2PO_4^-$

Based on the periodic table, as well as the table of ions, we may provide the chemical formulas of ionic compounds. We learnt how to do this in the section on "The periodic table". We also need to be able to provide the BALANCDED ionic equation for the formation of the ionic compound.

For example, for ammonium nitrate:
The ammonium ion is NH_4^+ and the nitrate ion is NO_3^-. It takes one ammonium ion and one nitrate ion to balance the charges. Therefore, $NH_4^+ + NO_3^- \rightarrow NH_4NO_3$

For lead(II) carbonate:
The lead ion is Pb^{2+} and the carbonate ion is CO_3^{2-}. It takes one lead ion and one carbonate ion to balance the charges. Therefore, $Pb^{2+} + CO_3^{2-} \rightarrow PbCO_3$

For sodium nitrate:
The sodium ion is Na^+ and the nitrate ion is NO_3^-. It takes one sodium ion and one nitrate ion to balance the charges. Therefore, $Na^+ + NO_3^- \rightarrow NaNO_3$

For sodium nitrite:
The sodium ion is Na^+ and the nitride ion is NO_2^-. It takes one sodium ion and one nitrite ion to balance the charges. Therefore, $Na^+ + NO_2^- \rightarrow NaNO_2$

For sodium nitride:
The sodium ion is Na^+ and the nitride ion is N^{3-}. It takes three sodium ions to balance out the charge of one nitride ion. Therefore, $3Na^+ + N^{3-} \rightarrow Na_3N$

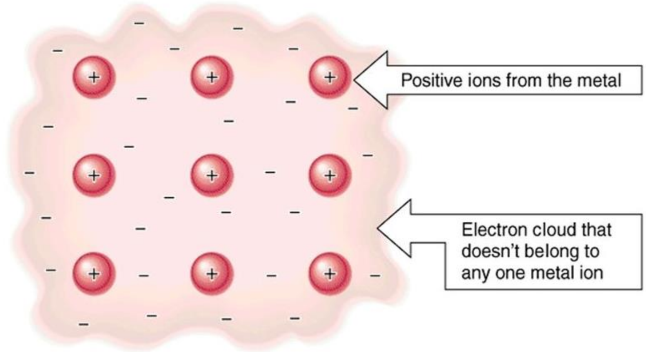
For tin(II) phosphate:
The tin(II) ion is Sn^{2+} and the phosphate ion is PO_4^{3-}. It takes three tin(II) ions to balance out the charge of two phosphate ions. Therefore, $3Sn^{2+} + 2PO_4^{3-} \rightarrow Sn_3(PO_4)_2$

For iron(III) chromate:
The iron(III) ion is Fe^{3+} and the chromate ion is CrO_4^{2-}. It takes two iron(III) ions to balance out the charge of two chromate ions. Therefore, $2Fe^{3+} + 3CrO_4^{2-} \rightarrow Fe_2(CrO_4)$

Metallic bonding

A metallic bond [D] is delocalised electrons that surround the positive metal ion core.

Most metals have many valence electrons. In a metallic bond, the metal atoms lose the strong attraction of their valence electrons because these electrons are equally attracted to the nuclei of other nearby metal atoms. The metal atoms become cations with a positive charge. The valence electrons from the metal atoms form a 'sea' of electrons that flow around the metal cations. The electrons may be described as delocalised electrons. See the diagram below for the electron-sea model of a metal showing positive metal cations in an ordered arrangement and a sea of negative electrons surrounding them.



Because metal cations and electrons have opposite charges, they are attracted to each other and also to other metal cations. These electrostatic forces are called metallic bonds and this holds the particles together to form a metallic network structure. This structure with a strong chemical bond gives metals their strength. The delocalised electrons allow metals to conduct electricity well.

Relative atomic mass and relative formula mass

We may calculate the **relative atomic mass (M)** of covalent molecules or **relative formula mass (M)** of ionic compounds. The calculation for M in both instances is the same. We use the relative atomic mass values represented on the periodic table, that is, the number below the element.

For example, the **relative atomic mass (M)** of NH_3 is:
M = 14+(3×1)
= 17 units
It is the relative *atomic* mass since NH_3 is covalently bonded. We add the atomic mass of N, which is 14, and the atomic mass of three H atoms, which is 1 for each H atom.

The **relative formula mass** of NaCl is:
M = 23+35.5
= 58.5 units
It is the relative *formula* mass since NaCl is ionically bonded. We add the atomic mass of Na, which is 23, and the atomic mass of Cl, which is 35.5.

Practice questions and answers

Practice questions

1. Represent each of the following molecules with Lewis dot diagrams using dots and crosses, and again using single, double or triple bonds:
   1.1 H_2
   1.2 CH_4
   1.3 NH_3
   1.4 HCl
   1.5 H_2O
   1.6 N_2
   1.7 O_2
   1.8 F_2
   1.9 Cl_2
   1.10 H_2S
   1.11 HF
   1.12 SiO_2
   1.13 CH_2O

2. Represent the half reactions, full reaction and final product of the following ionic compounds with Lewis dot diagrams using dots and crosses:
   2.1 $NaCl$
   2.2 AlN
   2.3 Al_2O_3

3. Complete the table to show what compounds the anions reflected in the rows would make with the cations reflected in the columns:

	K^+	Ca^{2+}	NH_4^+
OH^-			
O^{2-}			
NO_3^-			
PO_4^{3-}			

4. Provide the balanced ionic equation and the chemical formula of the following ionic compounds:

Iron(II) chromate	
Sodium oxide	
Barium sulfate	
Aluminium chloride	
Magnesium phosphate	
Tin(II) bromide	
Manganese(II) phosphide	
Iron(II) sulfate	
Zinc chloride	
Sodium carbonate	
Copper(II) nitrate	
Potassium permanganate	
Magnesium permanganate	
Manganese(IV) dichromate	
Manganese(II) hydrogen carbonate	
Chromium(III) hydrogen sulfate	

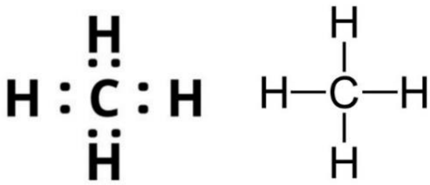
5. Calculate the relative atomic mass/relative formula mass of the following covalent molecules/ionic compounds. For each one, state whether you are calculating the relative atomic mass or relative formula mass:
   5.1 HCN
   5.2 CO_2
   5.3 Na_2CO_3
   5.4 NH_4OH
   5.5 $Cu(NO_3)_2$.

Practice question answers

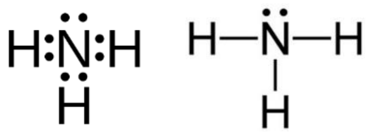
1.

1.1 H_2

1.2 CH_4



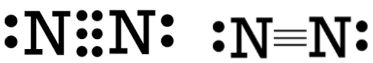
1.3 NH_3



1.4 HCl

1.5 H_2O

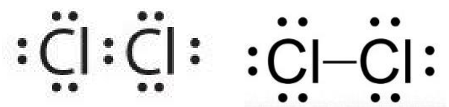
1.6 N_2



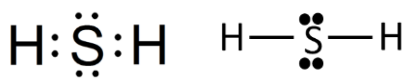
1.7 O_2

1.8 F_2

1.9 Cl_2

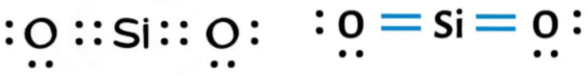


1.10 H_2S

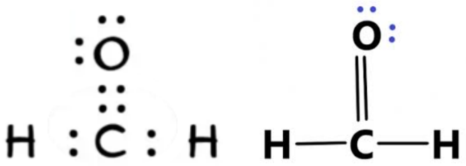


1.11 HF

1.12 SiO_2



1.13 CH_2O



2.

2.1 NaCl

Half reaction 1:

× Na → Na^+ + e-

Half reaction 2:

•• • Cl : •• + e- → ×• Cl : •• -

Full reaction:

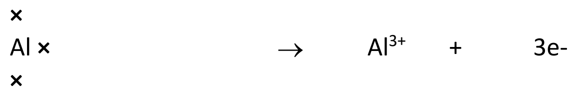
× Na + •• • Cl : •• → Na^+ + •• ×• Cl : •• -

Final product:

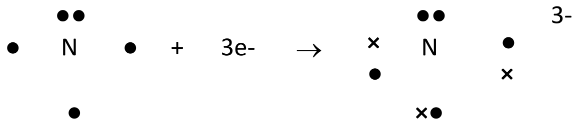
Na^+ [•• ×• Cl : ••] -

2.2 AlN

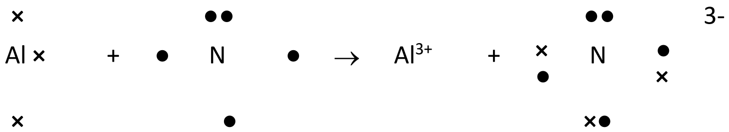
Half reaction 1:



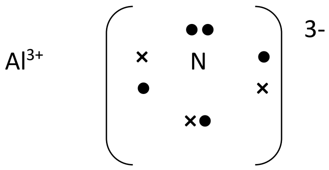
Half reaction 2:



Full reaction:

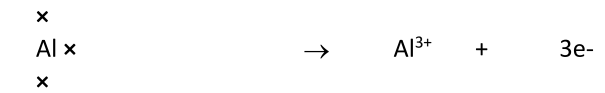


Final product:

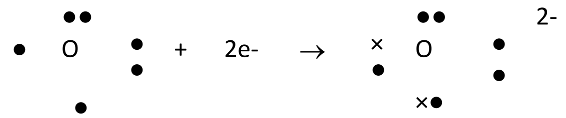


2.3 Al_2O_3

Half reaction 1:

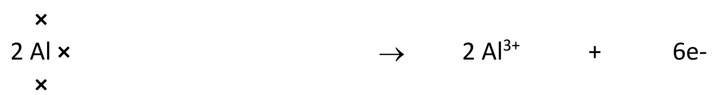


Half reaction 2:

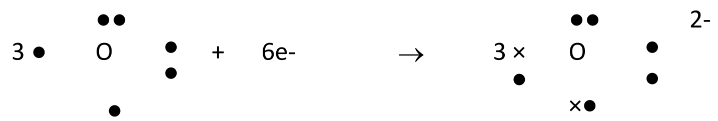


In this case, the number of electrons is not balanced. Aluminium can give away 3 electrons, but oxygen needs 2 electrons. We therefore need to balance out the number of electrons to get to the ratio in which the elements in this compound exist. This means that we need to balance out both half reactions. The first half reaction needs to be multiplied by 2 and the second half reaction needs to be multiplied by 3 to balance out the number of electrons in both equations:

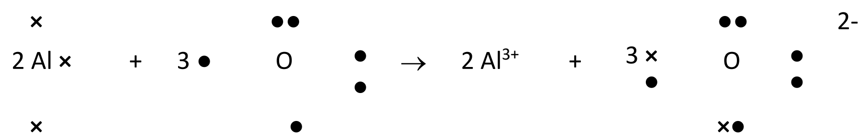
Half reaction 1 balanced:



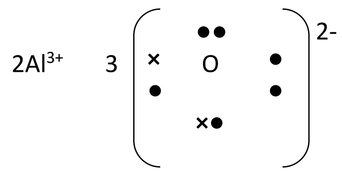
Half reaction 2 balanced:



Ful reaction:



Final product:



3.

	K^+	Ca^{2+}	NH_4^+
OH^-	KOH	$Ca(OH)_2$	NH_4OH
O^{2-}	K_2O	CaO	$(NH_4)_2O$
NO_3^-	KNO_3	$Ca(NO_3)_2$	NH_4NO_3
PO_4^{3-}	K_3PO_4	$Ca_3(PO_4)_2$	$(NH_4)_3PO_4$

4.

Iron(II) chromate	$Fe^{2+} + CrO_4^{2-} \rightarrow FeCrO_4$
Sodium oxide	$2Na^{+} + O^{2-} \rightarrow Na_2O$
Barium sulfate	$Ba^{2+} + SO_4^{2-} \rightarrow BaSO_4$
Aluminium chloride	$Al^{3+} + 3Cl^{-} \rightarrow AlCl_3$
Magnesium phosphate	$3Mg^{2+} + 2PO_4^{2-} \rightarrow Mg_3(PO_4)_2$
Tin(II) bromide	$Sn^{2+} + 2Br^{-} \rightarrow SnBr_2$
Manganese(II) phosphide	$3Mn^{2+} + 2P^{3-} \rightarrow Mn_3P_2$
Iron(II) sulfate	$Fe^{2+} + SO_4^{2-} \rightarrow FeSO_4$
Zinc chloride	$Zn^{2+} + 2Cl^{-} \rightarrow ZnCl_2$
Sodium carbonate	$2Na^{+} + CO_3^{2-} \rightarrow Na_2CO_3$
Copper(II) nitrate	$Cu^{2+} + 2NO_3^{-} \rightarrow Cu(NO_3)_2$
Potassium permanganate	$K^{+} + MnO_4^{-} \rightarrow KMnO_4$
Magnesium permanganate	$Mg^{2+} + 2MnO_4^{-} \rightarrow Mg(MnO_4)_2$
Manganese(IV) dichromate	$Mn^{4+} + 2Cr_2O_7^{2-} \rightarrow Mn(Cr_2O_7)_2$
Manganese(II) hydrogen carbonate	$Mn^{2+} + 2HCO_3^{-} \rightarrow Mn(HCO_3)_2$
Chromium(III) hydrogen sulfate	$Cr^{3+} + 3HSO_4^{2-} \rightarrow Cr(HSO_4)_3$

5.

5.1 HCN

Relative atomic mass (M) of HCN is:
M = 1+12+14
= 27 units

5.2 CO_2

Relative atomic mass (M) of CO_2 is:
M = 12+(16×2)
= 44 units

5.3 Na_2CO_3

Relative formula mass of Na_2CO_3:
M = (23×2)+12+(16×3)
= 106 units

5.4 NH_4OH

Relative formula mass of NH_4OH:
M = 14+(1×4)+16+1
= 35 units

5.5 $Cu(NO_3)_2$:

Relative formula mass of $Cu(NO_3)_2$:
M = 63,5+(14×2)+(16×6)
= 187,5 units

KEY/*SLEUTEL*

Electronegativity *Elektronegatiwiteit*	Atomic number *Atoomgetal*	Symbol *Simbool*	Approximate relative atomic mass *Benaderde relatiewe atoommassa*
1,9	29	Cu	63,5

1 (I)	2 (II)	3	4	5	6	7	8	9	10	11	12	13 (III)	14 (IV)	15 (V)	16 (VI)	17 (VII)	18 (VIII)
1 2,1 H 1																	2 He 4
3 1,0 Li 7	4 1,5 Be 9											5 2,0 B 11	6 2,5 C 12	7 3,0 N 14	8 3,5 O 16	9 4,0 F 19	10 Ne 20
11 0,9 Na 23	12 1,2 Mg 24											13 1,5 Aℓ 27	14 1,8 Si 28	15 2,1 P 31	16 2,5 S 32	17 3,0 Cℓ 35,5	18 Ar 40
19 0,8 K 39	20 1,0 Ca 40	21 1,3 Sc 45	22 1,5 Ti 48	23 1,6 V 51	24 1,6 Cr 52	25 1,5 Mn 55	26 1,8 Fe 56	27 1,8 Co 59	28 1,8 Ni 59	29 1,9 Cu 63,5	30 1,6 Zn 65	31 1,6 Ga 70	32 1,8 Ge 73	33 2,0 As 75	34 2,4 Se 79	35 2,8 Br 80	36 Kr 84
37 0,8 Rb 86	38 1,0 Sr 88	39 1,2 Y 89	40 1,4 Zr 91	41 Nb 92	42 1,8 Mo 96	43 1,9 Tc	44 2,2 Ru 101	45 2,2 Rh 103	46 2,2 Pd 106	47 1,9 Ag 108	48 1,7 Cd 112	49 1,7 In 115	50 1,8 Sn 119	51 1,9 Sb 122	52 2,1 Te 128	53 2,5 I 127	54 Xe 131
55 0,7 Cs 133	56 0,9 Ba 137	57 La 139	72 1,6 Hf 179	73 Ta 181	74 W 184	75 Re 186	76 Os 190	77 Ir 192	78 Pt 195	79 Au 197	80 Hg 201	81 1,8 Tℓ 204	82 1,8 Pb 207	83 1,9 Bi 209	84 2,0 Po	85 2,5 At	86 Rn
87 0,7 Fr	88 0,9 Ra 226	89 Ac															

58 Ce 140	59 Pr 141	60 Nd 144	61 Pm	62 Sm 150	63 Eu 152	64 Gd 157	65 Tb 159	66 Dy 163	67 Ho 165	68 Er 167	69 Tm 169	70 Yb 173	71 Lu 175
90 Th 232	91 Pa	92 U 238	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr