



Lesson 3: Ionic and molecular compounds (2.1 – 2.2)

Part 1: Properties of ionic compounds and molecular compounds

Table 4 Comparison of Physical Properties of Ionic and Molecular Compounds

Property	Ionic compound	Molecular compound
State at ambient temperature	crystalline solid	solid, liquid, or gas
Physical properties	hard, brittle	solids can be soft, waxy, flexible, or crystalline
Relative melting point/boiling point	high	low
Electrical conductivity when dissolved in water	good	poor (for most molecular compounds)
Electrical conductivity in the liquid state	good	poor
Examples	sodium chloride, calcium carbonate	water, carbon dioxide, methane, sucrose

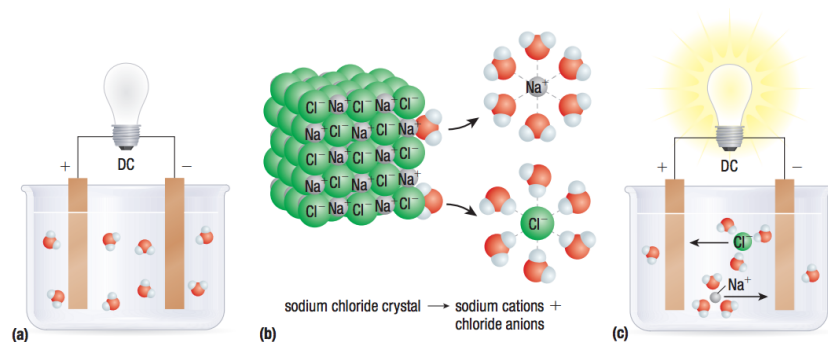


Figure 7 (a) Neutral water molecules (shown here as red and white models) do not carry charge. Therefore, they are poor conductors of electricity. (b) Sodium chloride dissolves in water, releasing sodium ions, Na^+ , and chloride ions, Cl^- . (c) The ions allow electricity to flow through the salt solution.



Figure 4 Sodium metal and chlorine gas react violently to form sodium chloride (table salt).

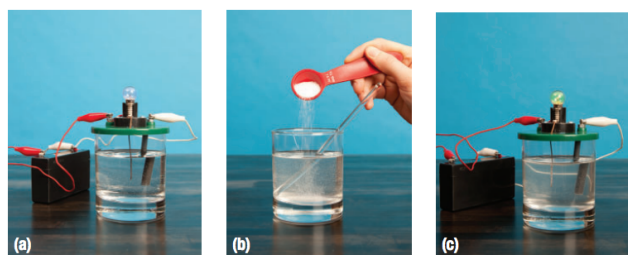


Figure 3 (a) Pure water is a poor conductor of electricity. (b) Sodium chloride dissolves in water to form a solution. (c) The resulting solution can conduct electricity extremely well. Sodium chloride is therefore an electrolyte.



Figure 2 (a) Carbon dioxide, (b) water, (c) paraffin wax, and (d) sugar are examples of molecular compounds.

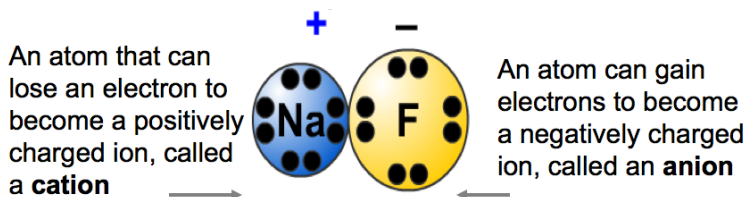


Part 2: The formation of ionic bonds

Ionic bonds are formed from the electrostatic attraction of positive and negative ions between metals and non-metals. Electrons are transferred.

Metallic elements will lose electrons to become cations with electron configurations of the previous noble gas.

Non-metal elements will gain electrons to become anions with electron configurations of the next noble gas.



Example: How does Magnesium combine with oxygen to form magnesium oxide?

And now does Aluminum combine with sulfur to form aluminum sulfide?



1 H 1+	(common ionic charges in compounds)																2 He n/a																		
	Nonmetals ↓ Metals																																		
3 Li 1+	4 Be 2+	Transition Metals										5 B n/a	6 C n/a	7 N 3-	8 O 2-	9 F 1-	10 Ne n/a																		
11 Na 1+	12 Mg 2+	13 Al 3+	14 Si n/a	15 P 3-	16 S 2-	17 Cl 1-	18 Ar n/a	19 K 1+	20 Ca 2+	21 Sc 3+	22 Ti 4+,3+,2+	23 V 5+,4+,3+,2+	24 Cr 6+,3+,2+	25 Mn 7+,4+,3+,2+	26 Fe 3+,2+	27 Co 3+,2+	28 Ni 3+,2+	29 Cu 2+,1+	30 Zn 2+	31 Ga 3+	32 Ge 4+,2+	33 As 5+,3+,3-	34 Se 2-	35 Br 1-	36 Kr n/a										
37 Rb 1+	38 Sr 2+	39 Y 3+	40 Zr 4+	41 Nb 5+,3+	42 Mo 6+	43 Tc 7+,6+,4+	44 Ru 3+	45 Rh 3+	46 Pd 4+,2+	47 Ag 1+	48 Cd 2+	49 In 3+	50 Sn 4+,2+	51 Sb 5+,3+,3-	52 Te 6+,4+,2-	53 I 1-	54 Xe n/a	55 Cs 1+	56 Ba 2+	57 La 3+	72 Hf 4+	73 Ta 5+	74 W 6+	75 Re 7+,6+,4+	76 Os 4+,3+	77 Ir 4+,3+	78 Pt 4+,2+	79 Au 3+,1+	80 Hg 2+,1+	81 Tl 3+,1+	82 Pb 4+,2+	83 Bi 5+,3+	84 Po 4+,2+,2-	85 At 1+,1-	86 Rn n/a
87 Fr 1+	88 Ra 2+	89 Ac 3+	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	Not Yet Named			113	114	115	116	117	118																		

How about ionic compounds containing transition metals?

For instance, Fe, Iron can hold either +3 or +2, two different charges to form Fe_2O_3 and $FeCl_2$, and Manganese can lose either 2 or 4 electrons to have 2+ and 4+ positive charges.

Example: Determine the possible charges of MnO_2 , K_3MnO_4 , K_2MnO_4 , $KMnO_4$, and Mn_2O_7 .



Part 3: The formation of covalent bonds

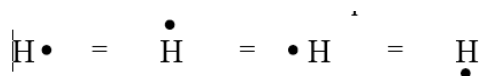
Covalent bonding occurs between two non-metals. Covalent bonding is different from ionic bonding because electrons are shared instead of transferred.

Drawing Lewis diagram/structure:

To draw Lewis Diagrams of an single atom:

Steps:

1. Write the element symbol. Around this, draw dots one for each valence electron.
2. The dots should be spread over four sides. Dots are not paired until all sides have at least one dot. It does not matter on which side dots are placed. For example, hydrogen can be drawn four ways:



To draw Lewis Diagrams of an molecular compounds:

Steps:

1. Place the central atom in the centre and other atoms around it evenly spaced.
2. Count the total number of valence electron for all atoms involved in the bonding.
3. Placed the electrons in pairs between the central atom and each non-central atom
4. Place the remaining electrons around the non-central atom until each has 8 electrons (H atoms have only 2 electrons.)
5. If electrons remain they are placed in pairs around the central atom.
6. Exception: if the central atom is in group 14, 15, 16, 17, 18, the octet rule must be satisfied by moving electron pairs from non-central atoms, creating multiple bonds. (Exceptions of octet rule called overfilled octet may apply)

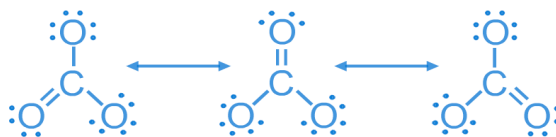


Example: Draw the Lewis diagrams of $AsBr_3$ and SO_2 .

Resonance structure:

When several structures with different electron distributions among the bonds are possible, all the structures contribute to the electronic structure of the molecule. These structures are called resonance structures

For instance: carbonate anion



Formal charge = Valence electron – NonBonding valence electrons – #of bonds

Carbon: 0	4	0	4
Oxygen:			
Oxygen:			



Summary/Hint:

When considering the resonance structure...

- Formal charge closest to zero is favoured
- Negative formal charges on the more negative elements are favoured.

When drawing Lewis diagrams...

- Symmetrical arrangements are favoured
- Hydrogen is never a central atom
- The **least** electronegative element is usually the central atom

Practice: Draw Lewis structure for the following compounds.

- 1) CCl_4
- 2) HCl
- 3) NH_3
- 4) OH^-
- 5) SO_4^{2-}
- 6) Nitrogen gas