

Ch. 4. Chemical naming

LL elements in the periodic table except for the noble gases—He, Ne, Ar, Kr, Xe, and Rn—combine to produce chemical compounds. Most of these chemicals are useful in your everyday life, and you drink water to quench your thirst, use Clorox to clean your house, or baking soda to get rid of a stinky refrigerator. In this chapter you will learn not only how to name these chemicals but also read chemical formulas—we call this to formulate chemicals. Still, chemical elements such as hydrogen and oxygen do not combine randomly and they only choose specific elemental partners to form a compound. As an example, hydrogen combines with oxygen using specific proportions to produce H₂O and not HO₂. In this chapter, you will also learn the rules that chemical elements use to combine.

4.1 lons & ionic charges

Atoms gain and lose electrons to produce ions. An ion is just an atom with a positive or negative charge. Ions result from an electron transfer. Positive ions have lost negatively charged electrons, whereas negative ions have gained electrons. The reason for this electron transfer is that atoms try to achieve a very stable electronic configuration with eight electrons in the valence, and this is called the octet electron configuration. Examples of ions are: H^+ , Ca^{2+} or O^{2-} . This section covers the properties of ions and ionic charges.

Cations and anions Atoms that lose electrons become positively charged. These ions are called cations. Examples of cations are Li⁺ or Mg²⁺ called lithium cation and magnesium cation, respectively. Atoms that gain electrons become negatively charged, as electrons have a negative charge. These ions are called anions. Examples of anions are F⁻ called fluoride or N³⁻ called nitride. The way to name anions is by using the name of the element and the suffix -ide.

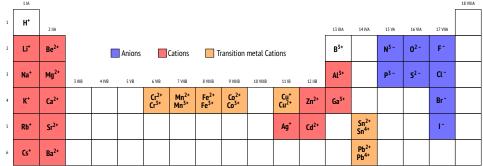


Figure 4.1 Ionic charges (valences) for different elements

Ionic charges: the valences How do we know that hydrogen produces a

H⁺ ion and nitrogen a N³⁻ anion? The charge of an ion is called an ionic charge, and the numbers are coming from the periodic table. H, Na, or K are in group IA (left of the table) and hence the ionic charge will be 1+. Similarly, Mg or Ca are in group IIA (left of the table) and hence the ionic charge will be 2+. Differently, F, Cl, or Br are in group 7A (right of the table) and their charge will be 1-. Oxygen is in group 6A (right of the table) and the ionic charge will be 2-. Figure 4.1 contains all ionic charges. What if the element is not on this list such as in the case of Iron (Fe)? In that case, very probably it will have several ionic charges and this charge has to be indicated in the chemical name. An example would be Fe, which ionic charge is not in Figure 4.1 as iron can have several ionic charges.

Sample Problem 1

Identify the correct ionic state of the following elements: (a) Cl (b) K (c) O (d) C

SOLUTION

Cl is on the 7A group and hence its charge is 1-, whereas potassium belongs to 1A and its charge will be 1+. Oxygen and carbon will have 2- and 4- charges. The final ionic states are: Cl^- , K^+ , O^{2-} and C^{4-} .

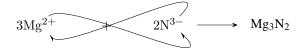
STUDY CHECK

Identify the correct ionic state of the following elements: (a) N (b) Br

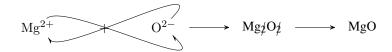
4.2 Ionic compounds

Ionic compounds are chemicals resulting from the combination of a nonmetallic element with a metallic element. An example is NaCl, which results from combining sodium (a metal) with chloride (a nonmetal). Ionic compounds normally have high melting points and are solid under normal conditions. A typical ionic compound would be NaCl, cooking salt. Atoms of an ionic compound are connected through an ionic bond. In an ionic bond, one element gives away electrons (the cation) and the other one receives electrons (the anion). As an example, in the NaCl molecule Na gives away an electron to Cl, and the molecule results from the combinations of Na⁺ and Cl⁺. In an ionic compound, the element on the left is positive and the one on the right is negative.

Combining ions Ionic compounds are the result of combining two ions: a positive (cation) and a negative (anion) ion. Each ion has a charge, depending on its location on the table. When combining two atoms you first need to arrange the ions starting from positive and followed by negative. The charges of an ion would become the coefficient of the other ion. For example Mg²⁺ and N³⁻ are combined as Mg₃N₂:



Another example would be the combination of Na^+ and O^{2-} that would be Na_2O . You need to simplify the indexes of the formula by dividing by the smallest one, always using integer values. For example, Mg^{2+} and O^{2-} give Mg_2O_2 that should be written as MgO



Another example that involves simplifying the formula is the chemical result of combining Ca²⁺ and C⁴⁻. After combining the charges we obtain Ca₄C₂ that needs to be simplified by dividing by the smallest number leading to Ca₂C.

Combine the following ions or give the ions given the final compound: (a) Li⁺ and O^{2-} (b) Ca^{2+} and O^{2-} (c) Li_3N (d) Mg_2C

SOLUTION

The result of combining Li⁺ and O²⁻ is Li₂O. For Ca²⁺ and O²⁻, the resulting chemical is CaO. Li₃N results from the combination of Li⁺ and N³⁻, and Mg₂C results from Mg²⁺ and C⁴⁻.

STUDY CHECK

Combine the following ions or give the ions given the final compound: (a) Na⁺ and F- (b) Na₃N

Simple ionic naming (type I ionic) Type I ionic compounds result from the combination of a metal with given valence (Li, Ca, Mg, etc.) and a nonmetal. To name an ionic compound (type I ionic) you need to (a) use the name of the first element in the compound, (b) use the first syllable of the second element, and (c) finish the name of the molecule in the suffix -ide. As an example, the formula NaCl is named sodium chloride, and MgCl₂ is named magnesium chloride. Another example would be:

calcium chloride CaCl₂ (ionic)

To formulate an ionic compound based on a name, we need to combine both ions by exchanging the valences (the ionic charges). For example, MgCl₂ results from the combination of Mg²⁺ and Cl⁻ so that the number 2 in MgCl₂ near the Cl atom is coming from the Mg²⁺. In other words:



The sign of the charges only indicates which element goes first in the formula: the positive element (cation) first followed by the negative element (anion). For example, the result of combining Na⁺ and Cl⁻ is NaCl and not ClNa as Na has a positive ionic charge and has to appear first in the formula.

Name or give the formula for the following ionic compounds: (a) MgO (b) Mg₃N₂ (c) Lithium nitride (d) Magnesium carbide

SOLUTION

The name for MgO is magnesium oxide. Mg₃N₂ is called magnesium nitride. The formula for Lithium nitride is Li₃N and the formula for Magnesium carbide is Mg₂C, result of simplifying Mg₄C₂ dividing by two, the smallest number.

STUDY CHECK

Name or give the formula for the following ionic compounds: (a) Sodium fluoride (b) Na₃N

The ionic chemical NaCl results from the combination of Na $^+$ and Cl $^-$. The ionic charges of Na and Cl are given in Figure 4.1 according to the group. If the ionic chemical contains a transition metal with variable ionic charge, that is, which is not in Figure 4.1 then the ionic naming becomes a bit more complex. The reason is that one needs to specify the charge of the metal, explicitly in the name of the chemical. An example would be NiCl $_2$ named as Nickel(II) chloride or Co $_2$ O $_3$ named as Cobalt(III) oxide.

Name complex ionic chemicals This section covers how to name ionic chemicals containing a metal with variable charge. In this case, you need to specify the charge of the metal in the name. To calculate this number you will solve a simple math equation. For example, the name of Mn₂O₃ is Manganese(III) oxide. How do we get this name? Manganese has several charges as it is not in Figure 4.1, let us use x for its charge Mn^x and oxygen has a charge of two O²⁻. After combining Mn^x and O²⁻ the resulting formula would be Mn₂Ox. By comparison with the given formula, Mn₂O₃, x has to be three and hence the charge of Mn has to be three. Therefore, the final name would be Manganese(III) oxide.

Properties of ionic compounds Ionic compounds normally have high melting points and are solid under normal conditions. A typical ionic compound would be NaCl, cooking salt.

The ionic bond Atoms of an ionic compound are connected through an ionic bond. In an ionic bond, one element gives away electrons (the cation) and the other one receives electrons (the anion). As an example, in the NaCl molecule Na gives away an electron to Cl, and the molecule results from the combinations of Na⁺ and Cl⁺. In an ionic compound, the element on the left is positive and the one on the right is negative.

Sample Problem 4

Name or give the formula for the following ionic compounds: (a) MnO (b) Fe_3N_2 (c) Cobalt(II) carbide (d) Iron(II) oxide

SOLUTION

All the chemicals on this example contain a metal that can have several charges, and hence, we need to specify the ionic charge on the name. MnO results from Mn^x and O^{2-} . After combining the ions, the formula would be Mn_2Ox , a formula that needs to be compared to MnO. The formulas do not look similar, so lets make them more similar by dividing by two so that $MnO\frac{x}{2}$ resembles MnO. By comparing x has to be 2 and hence the name is Manganese(II) oxide. The name for Fe_3N_2 would be Iron(II) nitride. The valence of Iron comes from combining Fe^x and N^{3-} that gives Fe_3Nx . By comparison with Fe_3N_2 x has to be two and the name is Iron(II) nitride. the formula for Cobalt(II) carbide would be Co_2C as Cobalt(II) is Co^{2+} and carbide is C^{4-} . After combining the ions one obtains Co_4C_2 that gives Co_2C . Finally, the formula for Iron(II) oxide is FeO as Iron(II) is Fe^{2+} and oxide is O^{2-} that gives Fe_2O_2 and simplifying one obtains FeO.

STUDY CHECK

Name or give the formula for the following ionic compounds: (a) Manganese(IV)

4.3 Covalent compounds

Covalent compounds are chemicals resulting from the combination of nonmetallic elements. And example is CO_2 , which results from combining carbon (a nonmetal) with oxygen (a nonmetal). At normal conditions, covalent compounds may exist as solids, liquids, or gases. Covalent compounds do not exhibit any electrical conductivity, either in pure form or when dissolved in water. A typical covalent compound would be H_2O , water. Atoms in a covalent compound are connected by means of a covalent chemical bond. In a covalent bond, both atoms connected share the electrons. As an example, the HCl molecule has an hydrogen and a chlorine atom connected by means of a covalent bond, in which each atoms share the electrons of the bond.

Covalent naming To name a covalent compound you need to (a) use the name of the first element in the compound, (b) use the first syllable of the second element, and (c) finish the name of the molecule in the suffix -ide. More importantly, you need to use prefixes that indicate the number of atoms in the molecule. See the Table below for a list of the different equivalencies between prefixes and numbers. As an example, the formula CH₄ is named carbon tetrahydride. Similarly, a covalent chemical name can be translated into a formula (we call this to formulate a chemical with a given name), and the formula for carbon monoxide would be CO. When the vowels a and o appear together, the first vowel is omitted as in carbon monoxide instead of carbon monoxide. Another example would be N₂O named as dinitrogen oxide, and the name sulfur hexafluoride corresponds to the formula SF₆. The prefix mono is omitted in the first element of the name, and for example, you will not name the chemical CO as monocarbon monoxide, you would just say carbon monoxide. A final example of a covalent compound:

dinitrogen pentoxide

 N_2O_5 (covalent)

Table 4.1 Prefixes used to name covalent compounds						
Prefix	number	Prefix	number	Prefix	number	
Mono	1	Tetra	4	Hepta	7	
Di	2	Penta	5	Octa	8	
Tri	3	Hexa	6	Nona	9	

Sample Problem 5

Name of give the name of the following covalent chemicals: (a) NO (b) CS_2 (c) Sulfur Dioxide (d) Nitrogen Trichloride

SOLUTION

All chemicals in this example are covalent as they result of the combination of nonmetals. In order to name them, we need to use prefixes and finish the sufix with -ide. The first chemical is called nitrogen monoxide. CS₂ is called carbon

disulfide. The formula for sulfur dioxide and nitrogen trichoride are respectively SO₂ and NCl₃.

STUDY CHECK

Name of give the name of the following covalent chemicals: (a) SCl₂ (b) diboron thrioxide

4.4 Naming acids & bases

In this section, we will learn how to name acids and bases. Acids normally have common names (e.g. sulfuric acid) and their naming does not follow modern rules. Names and formulas of acids are listed in tables. Differently, bases (e.g. sodium hydroxide) are named in a standard way.

Bases or hydroxides Bases (hydroxides) result from the combination of metal and the hydroxide anion (OH⁻). Examples are NaOH or Ca(OH)₂. The name of a base starts with the name of the cation finishing with the word *hydroxide*. An example is NaOH named as *sodium hydroxide*, or Ca(OH)₂, named as *calcium hydroxide*. The word *hydroxide* refers to the OH⁻ ion, and hence Sodium hydroxide results from combining Na⁺ and OH⁻, and Calcium hydroxide from combining Ca²⁺ and OH⁻. More examples of hydroxides:

Magnesium hydroxide

 $Mg(OH)_2(hydroxide)$

As a final note, not all bases are hydroxides. For example, ammonia NH₃ is a base even when it does not contain hydroxides in its structure. A way to remember this is sometimes to write ammonia like NH₄OH.

Acids Acids—in particular inorganic acids—are chemicals that normally contain hydrogen at the beginning of their formula. For example, HCl or H₂SO₄. HCl is a hydracid and is named as *hydrochloric acid*, whereas H₂SO₄ is an oxoacid that contains oxygen named as *sulfuric acid*. The names of acids are not standard and they come from common names employed in the field for many years. Table 4.1 contains a list of the most important oxoacids and hydracids. More examples of acids:

Nitric acid Hydrofluoric acid

HNO₃(oxoacid) HF(hydracid)

Sample Problem 6

Name or give the formula for the following acids and bases. Indicate whether the compound is an acid or a base. (a) HCN (b) KOH (c) Carbonic acid (d) Lithium hydroxide

SOLUTION

HCN is an acid named hydrocyanic acid. KOH is a base called potassium hydroxide. The formula for Carbonic acid is H_2CO_3 , and Lithium hydroxide is a base with formula LiOH.

STUDY CHECK

Name or give the formula for the following ionic compounds: (a) phosphoric acid (b) Mg(OH)₂

4.5 Oxidation states

Isolated atoms tend to have a neutral state. However, the atoms of elements have the capacity of gaining and losing electrons forming cations and anions. The atoms that form a compound can have different states resulting in losing and gaining electronic charge. We refer to this as the oxidation state of an element in a compound.

Oxidation states of oxoacids Consider the following set of acids: HClO, HClO₂, HClO₃ and HClO₄. We say Cl in these acids have different oxidation states or different oxidation numbers. This section will cover the calculation of the oxidation state of the central atom of an oxoacid.

Let us address the oxoacid: $HClO_3$. The goal is to calculate the oxidation number of the underlined element, Cl. To do this we will follow a set of simple rules. First, we will use the valences as the oxidation number of the elements to the right and the left of the central atom. Then, we will assign an unknown oxidation state of x to the central atom. After that, we will set up an equation so that the sum of all oxidation numbers equals the charge of the acid if any. In this formula, we will include the atomic coefficients. In the case of $HClO_3$, the equation would be:

$$1 + x + 3 \cdot (-2) = 0$$

as the number of oxygens is three, we will have to time by three the valence of oxygen. The number zero results from the charge of the acid. If we solve for x, we obtain x = 5. That is, the oxidation state of Cl on $H\underline{Cl}O_3$ is 5 and this is represented as $H\underline{Cl}^VO_3$.

Oxidizing and reducing character of oxoacids The importance of the oxidation state of the central elements of oxoacid results from the fact that acids with high or low oxidation states, tend to be very reactive, sometimes capable of completely dissolving metals. We call this oxidizing (or reducing) acids. For example, HNO3 and H₂SO₄ are both oxidizing acids, and these acids will solve for example a piece of copper. Similarly, acids with very small or negative oxidation numbers can be very reactive as well. These acids are called reducing acids or agents. Let us compare two oxoacids to elaborate more on the terminology used to describe redox numbers. For example, let us compare HCl^VO_3 and $HCl^{III}O_2$. We say Cl on HCl^VO_3 has a larger redox number than $HCl^{III}O_2$. We can also say, Cl in HCl^VO_3 is more oxidized than Cl on $HCl^{III}O_2$. Finally, we can also say, $HCl^{V}O_3$ is more reducing than $HCl^{III}O_2$. Again, the terms associated with high redox numbers are oxidized and reducing, and the terms associated with low redox numbers are reduced and oxidizing.

It is important to note that ultimately the oxidation state of an element is related to the number of electrons of the element. The more electrons the smaller-the more negative-the oxidation state. In other words, large oxidation states result from losing electrons.

Sample Problem 7

Calculate the redox number of S in the following acids and indicate the more oxidizing acid: $H_2S_2O_6$ named dithionic acid and $H_2S_2O_4$ named sulfuric acid.

We will set up the redox formula for the first acid $(H_2S_2O_6)$, given that the redox number of H is +1 and the redox number of O is -2.

$$2 \cdot 1 + 2 \cdot x + 6 \cdot (-2) = 0$$

Solving for x:

$$2 + 2 \cdot x - 12 = 0$$
 we have that $x = \frac{12 - 2}{2}$

The oxidation state of S in $H_2S_2O_3$ is +5. For the second acid (H_2SO_4) :

$$2 \cdot 1 + x + 4 \cdot (-2) = 0$$

Solving for x:

$$2 + x - 8 = 0$$
 we have that $x = \frac{8 - 2}{1}$

that gives a redox of 6. If we compare both acids the smaller the redox number the more reduced is the central element and the more oxidizing the acid is. Therefore, $H_2S_2O_3$ is more oxidizing than H_2SO_4 .

STUDY CHECK

Calculate the redox number of the following acids: (a) H₂MnO₄ (b) H₂Cr₂O₇

4.6 Naming complex salts & common chemicals

At this point, we saw the naming and formulation of ionic (e.g. NaCl) and covalent compounds (e.g. CO_2). This section covers the naming of complex salts: oxosalts and hydrosalts. In general, salts (oxosalts or hydrosalts) are the result of mixing an oxoacid and a base. They tend to look more complex than simple ionic or covalent compounds as they have at least three different elements. An example of oxosalt would be $CaSO_4$ called calcium carbonate. An example of hydrosalt would be $NaHSO_4$ which is called sodium monohydrosulfate. This section will also cover the naming of hydrates (e.g. $CaSO_4 \cdot H_2O$), which are compounds containing water molecules inside their structure. Before being able to name these complex chemicals it is convenient to practice combining ions, without paying attention to the naming.

Combining ions To combine two ions, you first arrange the positive ion on the left followed by the negative ion on the right, to then cross the ionic charges from the top of the ion to the bottom of the opposite ion. The positive and negative charges are not carried. If the ions have more than one element we have to use parenthesis. An example would be combining Ca²⁺ and PO₄³⁻ leading to Ca₃(PO₄)₂:

$$3Ca^{2+}$$
 + $2PO_4^{3-}$ $(a_3(PO_4)_2)$

We would simplify in case the charges compensate for each other.

An example would be combining Mg²⁺ and SO₄²⁻ leading to MgSO₄

$$\mathrm{Mg}^{2+}$$
 + SO_4^{2-} \longrightarrow $\mathrm{Mg}_{4}(\mathrm{SO}_{4})_{4}$ \longrightarrow $\mathrm{Mg}_{5}\mathrm{O}_{2}$

Table 4.2 Names of oxoacids and oxosalts (top table) and hyracids (bottom table).*					
Element	Oxoacid	Oxoacid Name	Oxoasalt	Oxoasalt Name	
Manganese	$HMnO_4$	Permanganic Acid	MnO_4^-	Permanganate	
	H_2MnO_4	Manganic acid	MnO_4^{-2}	Manganate	
Carbon	H_2CO_3	Carbonic Acid	CO_3^{-2}	Carbonate	
Nitrogen	HNO_3	Nitric Acid	NO ₃ -	Nitrate	
	HNO_2	Nitrous Acid	NO_2^-	Nitrite	
Phosphorus	H_3PO_4	Phosphoric Acid	PO_4^{-3}	Phosphate	
Sulfur	H_2SO_4	Sulfuric Acid	SO_4^{-2}	Sulfate	
	H_2SO_3	Sulfurous Acid	SO_3^{-2}	Sulfite	
	$H_2S_2O_2$	Thiosulfurous Acid	$S_2O_2^{-2}$	Thiosulfite	
	$H_2S_2O_3$	Thiosulfuric Acid	$S_2O_3^{-2}$	Thiosulfate	
	$H_2S_2O_7$	Disulfuric acid	$S_2O_7^{-2}$	Disulfate	
	$H_2S_2O_8$	Peroxodisulfuric acid	$S_2O_8^{-2}$	Peroxodisulfate	
Chlorine	HClO ₄	Perchloric Acid	ClO ₄ -	Perchlorate	
	HClO ₃	Chloric acid	ClO ₃	Chlorate	
	HClO ₂	Chlorous acid	ClO ₂	Chlorite	
	HClO	Hypochlorous acid	ClO-	Hypochlorite	
Iodine	HIO_4	Periodic Acid	IO_4	Periodate	
Chromium	H_2CrO_4	Chromic acid	CrO_4^{2-}	Chromate	
	$H_2Cr_2O_7$	Dichromic acid	$Cr_2O_7^{2-}$	Dichromate	
Boron	H_3BO_3	Boric acid	BO ₃ ³⁻	Borate	
Hydracid	Hydracid Name	Hydracid	Hydracid Name		
HCl	Hydrochloric acid	HBr	Hydrobromic acid		
HI	Hydroiodic acid	HF	Hydrofluoric acid		
HCN	Hydrocyanic acid	H_2S	Hydrosulfuric acid		

^{*} Yellow indicate very important acids

Naming Oxosalts The names of the oxosalts are constructed by combining the name of the first element-you need to specify its charge in the case of a transition metal element with different possible charges-followed by the name of the oxosalt from Table 4.2. For example, the name of MgSO₄ is magnesium sulfate, as Mg²⁺ is magnesium and SO_4^{2-} is sulfate. Another example is $Fe_2(CO_3)_3$ called Iron(III) carbonate. A final example:

Litium nitrate	LiNO ₃ (oxosalt)		

Formulating Oxosalts In the case that you know the name of an oxosalt and you want to obtain its formula, you first need to arrange the positive ion on the left followed by the negative ion on the right, to then cross the ionic charges from the top of the ion to the bottom of the opposite ion. For example, calcium nitrate results from the combination of Ca^{2+} calcium and NO_3^- , nitrate. By combining the two ions we obtain the final formula as $Ca(NO_3)_2$:

$$\operatorname{Ca}^{2+}$$
 + $\operatorname{NO}_3^ \operatorname{Ca}(\operatorname{NO}_3)_2$

Sample Problem 8

Name of give the name of the following oxosalts: (a) K_2SO_4 (b) Na_2CO_3 (c) Magnesium carbonate (d) Sodium phosphate

SOLUTION

 K_2SO_4 is named potassium sulfate, as K^+ is potassium and SO_4^{2-} stands for sulfate. Na_2CO_3 is sodium carbonate. Magnesium carbonate is $MgCO_3$ and sodium phosphate is Na_3PO_4 .

STUDY CHECK

Name of give the name of the following oxosalts: (a) $CaSO_4$ (b) Aluminum sulfate

Some oxosalts contain hydrogen atoms in their structure between the oxosalt cation and anion (e.g. NaHSO₄). For example, NaHSO₄ is named sodium monohydrogen-sulfate. To understand this name, we will first focus on the second part of the name, monohydrogensulfate which represents the anion. The name monohydrogensulfate (HSO₄ $^-$) comes from adding a proton (H⁺) to a sulfate cation (SO₄ 2 $^-$). Mind that protons (H⁺) are positively charged and therefore if we add a single H⁺ to a sulfate cation (SO₄ 2 $^-$) the charge will have to decrease a single unit, giving us HSO₄ $^-$. As we can see, the name is directly related to the oxosalt anion and the number of hydrogens in the hydrosalt name. For example, phosphate (PO₄ 3 $^-$) is an oxosalt anion whereas hydrogenphosphate (HPO₄ 2 $^-$) and dihydrogenphosphate (H2PO₄ $^-$) are both hydrosalt anions. An explanation about the charges: as phosphate has three negative charges, hydrogenphosphate has to have one less charge (that is 2 $^-$) and dihydrogenphosphate has to have two less negative charges (that is $^-$ 1). Some final hydrosalt anions examples:

carbonate ${\rm CO_3}^{-2}({\rm oxosalt\ ion})$ monohydrogen carbonate ${\rm HCO_3}^{-}({\rm hydrosalt\ ion})$

Above we saw how to name just the ending of the oxosalt with hydrogen anion. Now we can move forward to the whole naming of the salt. We just need to add the name of the element in the first place, and for example, NaH₂BO₃ would be named sodium dihydrogenborate. If the first ion—the cation—is a transition metal cation (a type two cation) we need to include in parenthesis the valence of the cation. For example, Fe(H₂BO₃)₂ would be named iron(II) dihydrogenborate. More examples:

sodium carbonate Na₂CO₃(oxosalt) sodium monohydrogen carbonate NaHCO₃ (hydrosalt)

Sample Problem 9

Name or formulate the following hydrosalts: (a) Magnesium hydrogensulfate (b) Sodium hydrogen carbonate (c) LiHCO $_3$ (d) Mg(H $_2$ PO $_4$) $_2$

SOLUTION

The formula of Magnesium hydrogensulfate is $Mg(HSO_4)_2$ as the formula for monohydrogen sulfate is HSO_4^- and the valence of magnesium is Mg^{2+} . The formula for Sodium monohydrogen carbonate is NaHCO₃ as it results from combining Na⁺ and HCO₃⁻. Mind monohydrogen carbonate results from adding a hydrogen ion H⁺ to a carbonate CO_3^{2-} ion. The name for LiHCO₃ is lithium monohydrogen carbonate, whereas the name for $Mg(H_2PO_4)_2$ is magnesium dihydrogenphosphate.

STUDY CHECK

Name or formulate the following hydrogensalts: (a) LiHS₂O₃ (b) LiH₂PO₄ (c) sodium hydrogenphosphate

Naming oxosalts with hydrogeHydrates Some chemicals contain water molecules trapped in their structure and therefore water molecules (H₂O) are often indicated in chemical formulas. These types of chemicals containing water are called hydrates, precisely because hydrate means water. Examples of hydrates are BeSO₄ · 4 H₂O or CuSO₄ · 5 H₂O called respectively beryllium sulfate tetrahydrate and cupper(II) sulfate pentahydrate. To formulate hydrates you just need to use prefixes such as mono, di, tetra—the same ones we use to name covalent chemicals—to indicate the number of water molecules in the chemical and end the name with hydrate. As a note, warming up hydrates (e.g. BeSO₄ · 4 H₂O) results in the release of water producing a dehydrated or anhydrous compound (e.g. BeSO₄). A final example of hydrate naming:

Sodium sulfate pentahydrate

 $Na_2SO_4\cdot 5\,H_2O\;(a\;\text{hydrate})$

Sample Problem 10

Name or formulate the following hydrates: (a) Nickel(II) permanganate dihydrate (b) Sodium nitrate monohydrate (c) Na $_2$ CO $_3 \cdot 10\,H_2$ O (d) MgSO $_4 \cdot 7\,H_2$ O SOLUTION

The formula for Nickel(II) permanganate is Ni(MnO₄)₂, therefore the formula for Nickel(II) permanganate dihydrate is Ni(MnO₄)₂ · 2 H₂O. The formula for Sodium nitrate is NaNO₃, therefore NaNO₃ · H₂O is Sodium nitrate monohydrate. The name for Na₂CO₃ · 10 H₂O is sodium carbonate decahydrate and MgSO₄ · 7 H₂O is magnesium sulfate heptahydrate.

STUDY CHECK

Name or formulate the following hydrates: (a) LiNO $_3 \cdot H_2O$ (b) Na $_3PO_4 \cdot 3 H_2O$ (c) sodium sulfate tetrahydrate

Common naming Some of the chemicals are normally referred to by a common name that does not involve the use of any chemical naming rules. An example would be H₂O normally referred to as water instead of its standard name which is dihydrogen oxide. You can find more names in Table 4.2. Another example:

NaCl Sodium chloride (standard name) Table salt (common name)

Table 4.3 List of common chemicals						
Chemical	Name	Chemical	Name			
H ₂ O	Water	Mg(OH) ₂	Milk of magnesia			
NH_3	Ammonia	N_2O	Laughing gas			
CH_4	Methane	CaCO ₃	Marble			
CO_2	Dry ice	CaO	Quicklime			
NaCl	Table salt	$NaHCO_3$	Baking Soda			
$NaHCO_3$	Sodium Bicarbonate	$MgSO_4 \cdot 7 H_2O$	Epsom Salt			

Sample Problem 11

Name or formulate the following common chemicals: milk of magnesia and dry ice.

SOLUTION

The formula for milk of magnesia is $Mg(OH)_2$ (magnesium hydroxide), whereas dry ice is the common name for CO_2 , carbon dioxide.

STUDY CHECK

Name or formulate the following common chemicals: (a) ammonia (b) methane