Ch. 4. Chemical naming

ALL 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 H_2O and not HO_2. In this chapter, you will also learn the rules that chemical elements use to combine.

4.1 Ions & 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 Mg^{2+} 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 N^{3-} called nitride. The way to name anions is by using the name of the element and the suffix -ide.

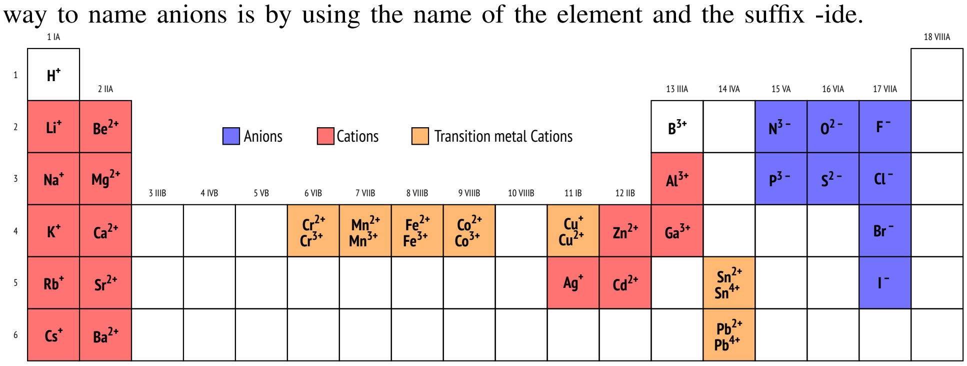

Figure 4.1 Ionic charges (valences) for different elements

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

H^+ ion and nitrogen a N^{3-} 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 Mg^{2+} and N^{3-} are combined as Mg_3N_2:

$$3Mg^{2+} \qquad 2N^{3-} \longrightarrow Mg_3N_2$$

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

$$Mg^{2+} \quad O^{2-} \longrightarrow Mg_{\cancel{2}}O_{\cancel{2}} \longrightarrow MgO$$

Another example that involves simplifying the formula is the chemical result of combining Ca^{2+} and C^{4-}. After combining the charges we obtain Ca_4C_2 that needs to be simplified by dividing by the smallest number leading to Ca_2C.

Sample Problem 2

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 O^{2-} is Li_2O. For Ca^{2+} and O^{2-}, the resulting chemical is CaO. Li_3N results from the combination of Li^+ and N^{3-}, and Mg_2C results from Mg^{2+} and C^{4-}.

STUDY CHECK

Combine the following ions or give the ions given the final compound: (a) Na^+ and F^- (b) Na_3N

`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 $MgCl_2$ is named magnesium chloride. Another example would be:

calcium chloride	$CaCl_2$ (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,$MgCl_2$ results from the combination of Mg^{2+} and Cl^- so that the number 2 in $MgCl_2$ near the Cl atom is coming from the Mg^{2+}. In other words:

$$Mg^{2+} \quad + \quad 2Cl^- \longrightarrow MgCl_2$$

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.

Sample Problem 3

Name or give the formula for the following ionic compounds: (a) MgO (b) Mg_3N_2 (c) Lithium nitride (d) Magnesium carbide

SOLUTION

The name for MgO is magnesium oxide. Mg_3N_2 is called magnesium nitride. The formula for Lithium nitride is Li_3N and the formula for Magnesium carbide is Mg_2C, result of simplifying Mg_4C_2 dividing by two, the smallest number.

STUDY CHECK

> Name or give the formula for the following ionic compounds: (a) Sodium fluoride (b) Na_3N

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_2O_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 Mn_2O_3 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 O^{2-}. After combining Mn^x and O^{2-} the resulting formula would be Mn_2Ox. By comparison with the given formula, Mn_2O_3, 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) oxide (b) $AuCl_3$

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 CH_4 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 ~~monooxide~~. Another example would be N_2O named as dinitrogen oxide, and the name sulfur hexafluoride corresponds to the formula SF_6. 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. CS_2 is called carbon disulfide. The formula for sulfur dioxide and nitrogen trichoride are respectively SO_2 and NCl_3.

STUDY CHECK

Name of give the name of the following covalent chemicals: (a) SCl_2 (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 Ca^{2+} and OH^-. More examples of hydroxides:

Magnesium hydroxide	$Mg(OH)_2$(hydroxide)

As a final note, not all bases are hydroxides. For example, ammonia NH_3 is a base even when it does not contain hydroxides in its structure. A way to remember this is sometimes to write ammonia like NH_4OH.

`Acids` Acids–in particular inorganic acids–are chemicals that normally contain hydrogen at the beginning of their formula. For example, HCl or H_2SO_4. HCl is a hydracid and is named as *hydrochloric acid*, whereas H_2SO_4 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	HNO_3(oxoacid)
Hydrofluoric acid	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, $HClO_2$, $HClO_3$ and $HClO_4$. 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: $H\underline{Cl}O_3$. 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 $H\underline{Cl}O_3$, 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$H\underline{Cl}^{V}O_3$.

`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, HNO_3 and H_2SO_4 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 $H\underline{Cl}^{V}O_3$ and $H\underline{Cl}^{III}O_2$. We say Cl on $H\underline{Cl}^{V}O_3$ has a larger redox number than $H\underline{Cl}^{III}O_2$. We can also say, Cl in $H\underline{Cl}^{V}O_3$ is more oxidized than Cl on $H\underline{Cl}^{III}O_2$. Finally, we can also say, $H\underline{Cl}^{V}O_3$ is more reducing than $H\underline{Cl}^{III}O_2$. 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: $H_2\underline{S}_2O_6$ named dithionic acid and $H_2\underline{S}O_4$ named sulfuric acid.

SOLUTION

We will set up the redox formula for the first acid ($H_2\underline{S}_2O_6$), 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:

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

The oxidation state of S in $H_2\underline{S}_2O_3$ is $+5$. For the second acid ($H_2\underline{S}O_4$):

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

Solving for x:

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

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, $H_2\underline{S}_2O_3$ is more oxidizing than $H_2\underline{S}O_4$.

STUDY CHECK

Calculate the redox number of the following acids: (a) H_2MnO_4 (b) $H_2Cr_2O_7$

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 Ca^{2+} and ${PO_4}^{3-}$ leading to $Ca_3(PO_4)_2$:

$$3Ca^{2+} \qquad + \qquad 2PO_4^{3-} \longrightarrow Ca_3(PO_4)_2$$

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

An example would be combining Mg^{2+} and $SO_4{}^{2-}$ leading to $MgSO_4$

$$Mg^{2+} \quad + \quad SO_4^{2-} \longrightarrow Mg_{\cancel{2}}(SO_4)_{\cancel{2}} \longrightarrow MgSO_4$$

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 $MgSO_4$ is magnesium sulfate, as Mg^{2+} 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 $Ca(NO_3)_2$:

$$Ca^{2+} \quad + \quad NO_3^- \longrightarrow Ca(NO_3)_2$$

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. $NaHSO_4$). For example, $NaHSO_4$ is named sodium monohydrogensulfate. To understand this name, we will first focus on the second part of the name, monohydrogensulfate which represents the anion. The name monohydrogensulfate ($HSO_4{}^-$) comes from adding a proton (H^+) to a sulfate cation ($SO_4{}^{2-}$). Mind that protons (H^+) are positively charged and therefore if we add a single H^+ to a sulfate cation ($SO_4{}^{2-}$) the charge will have to decrease a single unit, giving us $HSO_4{}^-$. 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 ($PO_4{}^{3-}$) is an oxosalt anion whereas hydrogenphosphate ($HPO_4{}^{2-}$) and dihydrogenphosphate ($H_2PO_4{}^-$) 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	$CO_3{}^{-2}$(oxosalt ion)
monohydrogen carbonate	$HCO_3{}^-$(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, NaH_2BO_3 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, $Fe(H_2BO_3)_2$ would be named iron(II) dihydrogenborate. More examples:

sodium carbonate	Na_2CO_3(oxosalt)
sodium monohydrogen carbonate	$NaHCO_3$ (hydrosalt)

Sample Problem 9

Name or formulate the following hydrosalts: (a) Magnesium hydrogensulfate (b) Sodium hydrogen carbonate (c) $LiHCO_3$ (d) $Mg(H_2PO_4)_2$

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 $NaHCO_3$ as it results from combining Na^+ and HCO_3^-. Mind monohydrogen carbonate results from adding a hydrogen ion H^+ to a carbonate CO_3^{2-} ion. The name for $LiHCO_3$ 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) $LiHS_2O_3$ (b) LiH_2PO_4 (c) sodium hydrogenphosphate

`Naming oxosalts with hydrogeHydrates` Some chemicals contain water molecules trapped in their structure and therefore water molecules (H_2O) are often indicated in chemical formulas. These types of chemicals containing water are called *hydrates*, precisely because hydrate means water. Examples of hydrates are $BeSO_4 \cdot 4\,H_2O$ or $CuSO_4 \cdot 5\,H_2O$ 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. $BeSO_4 \cdot 4\,H_2O$) results in the release of water producing a dehydrated or *anhydrous* compound (e.g. $BeSO_4$). A final example of hydrate naming:

Sodium sulfate pentahydrate	$Na_2SO_4 \cdot 5\,H_2O$ (a hydrate)

Sample Problem 10

Name or formulate the following hydrates: (a) Nickel(II) permanganate dihydrate (b) Sodium nitrate monohydrate (c) $Na_2CO_3 \cdot 10\,H_2O$ (d) $MgSO_4 \cdot 7\,H_2O$

SOLUTION

The formula for Nickel(II) permanganate is $Ni(MnO_4)_2$, therefore the formula for Nickel(II) permanganate dihydrate is $Ni(MnO_4)_2 \cdot 2\,H_2O$. The formula for Sodium nitrate is $NaNO_3$, therefore $NaNO_3 \cdot H_2O$ is Sodium nitrate monohydrate. The name for $Na_2CO_3 \cdot 10\,H_2O$ is sodium carbonate decahydrate and $MgSO_4 \cdot 7\,H_2O$ 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 H_2O 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_2O	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