Chapter 7
How Do We Name Chemical Compounds?

In conversation, chemists do not use chemical formulas, but chemical names. It is important to see how the naming process works and to be able to write names corresponding to chemical formulas and the reverse.

7.1 Background

When the first chemical compounds were identified or synthesized, a variety of ways were used to name them. Some were named after their taste, such as ‘sugar of lead.’ Others were named after the place where they were first discovered, such as ‘Epsom salts.’ Yet others were named after their discoverer, such as ‘Magnus’s green salt.’

It was in 1782 that the French chemist Louis-Bernard Guyton de Morveau devised a logical system of naming which involved the name of the key constituent element combined with suffixes and prefixes to uniquely identify the compound. By the early twentieth century, more compounds had been discovered than de Morveau’s system could handle. It was a German chemist, Alfred Stock, who, in 1934, proposed a modification of de Morveau’s system and it is this hybrid method that we use today. Even then, from year-to-year there are small alterations that need to be made. This task is the responsibility of the International Union of Pure and Applied Chemistry (IUPAC). Every few years, IUPAC issues a book (Figure 7.1) – and now a website – on updates in the naming system.

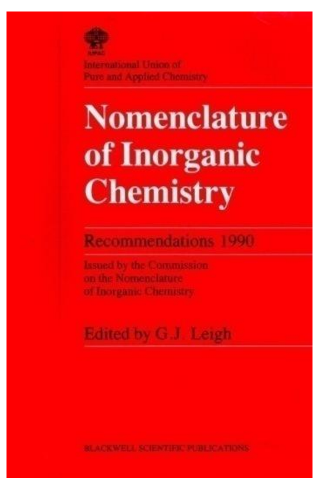


Figure 7.1 The IUPAC “Red Book” – the definitive guide to the naming of inorganic compounds.

7.2 An Overview of Naming Chemical Compounds

There are actually two different categories of compounds, each of which has its own rules. There are organic compounds, which are mainly or exclusively compounds of carbon and hydrogen (often with oxygen, as well) and these compounds, and their naming system, will be covered in Chapters 15 and 16. Then there are inorganic compounds which are combinations of all the other elements. The naming of these compounds depends upon whether they are ionic compounds or covalent compounds. A summary is shown in Figure 7.2.

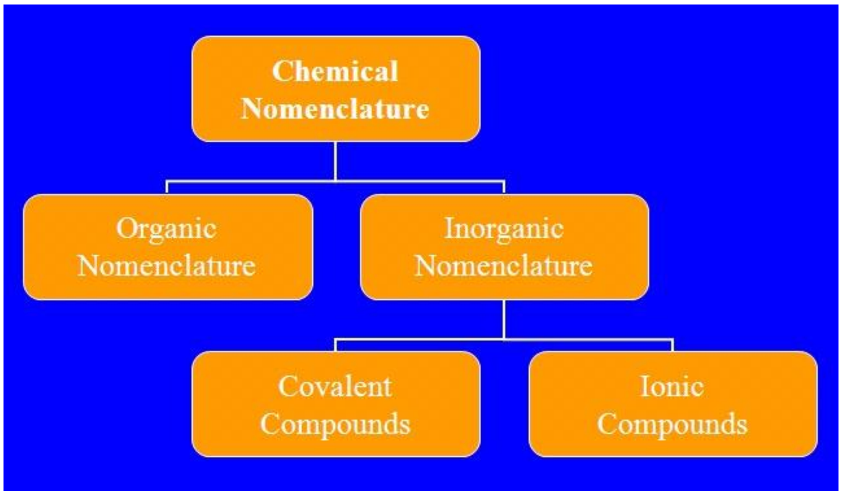


Figure 7.2 A classification of the modern naming systems for chemical compounds.

7.3 The Charges and Names of Common Ions

In Chapter 6, Sections 6.2 and 6.3, it was shown how compounds could form by the transfer of electrons between atoms. The element with the fewer outer electrons would lose these electrons to give a positive ion (the cation). The element with the greater number of outer electrons, would gain more electrons to give a negative ion with a full outer layer (the anion). In this Section, the common ions will be listed, together with their corresponding names.

THE NAMING OF CATIONS

For most main-group metals, there is only one possible charge on the cation – that corresponding to a completely empty outer electron layer. The name of the cation is taken as the same as the name of the element itself. In a few cases, a metal can form cations with two different charges. To distinguish the two possibilities, the charge on the ion is indicated by Roman numerals in parentheses. For example, lead (in Group 14) can lose all four of its outer electrons to form an ion of charge +4, but it can also lose only two electrons to form an ion of charge +2. To differentiate the two ions, they would be named lead(IV) ion and lead(II) ion. The Arabic numerals and their corresponding Roman numerals are shown in Table 7.1.

Table 7.1 The Roman numerals corresponding to the first ten Arabic numerals.

Arabic	1	2	3	4	5	6	7	8	9	10
Roman	I	II	III	IV	V	VI	VII	VIII	IX	X

THE NAMING OF ANIONS

To name an anion, the ending of the element name is removed and the ending (suffix) of –*ide* is added. For example, the element sulfur with six outer electrons will add two more. The resulting anion, S^{2-}, will be called sulfide.

PATTERNS AMONG THE IONS

As was discussed in Chapter 6, Sections 6.2 and 6.3, there are patterns among the ions that are formed. Thus an atom of each element in Group 1 possesses one outer electron and will lose that to form an ion of charge +1. Similarly, an atom of each element in Group 17 possesses seven electrons in its outer layer and will gain one electron to form an ion of charge −1. For the important main group elements, the charges and the corresponding names are listed in Table 7.2.

Table 7.2 The charges and names of the common ions of the main group elements.

Group 1	Group 2		Group 13	Group 14	Group 15	Group 16	Group 17
Li^{+} lithium					N^{3-} nitride	O^{2-} oxide	F^{-} fluoride
Na^{+} sodium	Mg^{2+} magnesium		Al^{3+} aluminum		P^{3-} phosphide	S^{2-} sulfide	Cl^{-} chloride
K^{+} potassium	Ca^{2+} calcium						Br^{-} bromide
Rb^{+} rubidium	Sr^{2+} strontium			Sn^{2+}, Sn^{4+} tin(II), tin(IV)			I^{-} iodide
Cs^{+} cesium	Ba^{2+} barium			Pb^{2+}, Pb^{4+} lead(II), lead(IV)			

There are also a few transition metals for which it is important to know the charges on their ions and their corresponding names. Unfortunately, for the transition metals, there is not a simple link between the Group number and charge.

Table 7.3 The charges and names of the common ions of the transition metals.

Group 8	Group 9	Group 10	Group 11	Group 12
Fe^{2+}, Fe^{3+} iron(II), iron(III)	Co^{2+} cobalt(II)	Ni^{2+} nickel(II)	Cu^{+}, Cu^{2+} copper(I), copper(II)	Zn^{2+} zinc
				Cd^{2+} cadmium
				Hg^{+}, Hg^{2+} mercury(I). mercury(II)

7.4 The Formulas and Names of Binary Ionic Compounds

The previous section introduced the formulas and names of the common ions and this section will show how the formulas of ionic compounds consisting of two elements (binary compounds) can be constructed without the necessity of drawing electron-dot symbols. Along with determining the formulas, the corresponding names will be derived.

FORMULAS OF IONIC COMPOUNDS

The formation of ionic compounds is explained in terms of the following two concepts:

1. Positive ions (cations) will combine with negative ions (anions)
2. Cations and anions will only combine in ratios that will result in electrically-neutral compounds.

For example, to find the formula of the compound between barium and chlorine, it is first necessary to write the ions with their charges: Ba^{2+} and Cl^{-}. To balance the charges, two anions will be needed for each cation. Thus the formula will be $BaCl_2$.

NAMING IONIC COMPOUNDS

When writing about a compound, it is the name of a compound that is usually given. The rules of naming binary ionic compounds are as follows:

1. The name consists of two words
2. The name of the metal is placed first
3. The name of the nonmetal has the ending *–ide*
4. If the metal ion can have more than one charge, then that charge is written as Roman numerals in parenthesis.

To illustrate, the name of $BaCl_2$ can be found as follows using Rules 1-3. The name for the Ba^{2+} is "barium" and the name for Cl^{-} is "chloride". So the name will be barium chloride.

If it is a cation which could have more than one charge, then the process is slightly more convoluted and involves Rule 4. For example, to find the name of $PbCl_4$, it is necessary to recall that lead can form an ion of charge +2 and one of +4. Each chloride ion has a charge of −1 and as there are four chloride ions, there is a total negative charge of −4. To balance, the cation must have a charge of +4. The name of the compound would therefore be lead(IV) chloride.

Sometimes it is necessary to derive the corresponding formula, given the name. As an example, to find the formula of sodium nitride, it is first necessary to write the formulas of the constituent ions: Na^{+} and N^{3-}. Three Na^{+} cations are necessary to balance the charge of one N^{3-} anion, so the formula will be Na_3N.

Example 7.1

Determine the name of (a) SnO and (b) SnO_2.

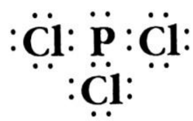
Answer

(a) The oxide ion always has a charge of –2. So the tin ion must have a charge of +2. The name is tin(II) oxide.

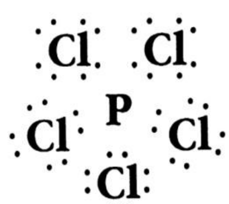
(b) The oxide ion always has a charge of –2, and as there are two, the total charge is –4. So the tin ion must have a charge of +4. The name is tin(IV) oxide.

7.5 The Formulas and Names of Binary Covalent Compounds

The formulas for covalent compounds are not predictable in a beginning chemistry course. In Chapter 6, Section 6.4, it was shown how pairs of nonmetal atoms combine to form covalent-bonded molecules. Normally, each atom (except hydrogen) attempts to acquire eight electrons in its outer layer – sometimes referred to as an *octet*. For example, a phosphorus atom with five outer electrons combines with three chlorine atoms, each one of which has seven electrons to give PCl_3 as shown in the electron-dot structure below.



For elements of the 2nd Period, the ‘central’ atom can never exceed a maximum of eight electrons. However, for elements of the 3rd and subsequent periods, the central atom can have ten or twelve electrons. For example, from the electron-dot figure above, it is possible for the lone pair of electrons to split and each one being shared with another chlorine atom. The electron-dot structure for the new compound, PCl_5, will be:



For binary covalent compounds, a different naming system is used, one that uses prefixes as numbers. The rules of naming binary covalent compounds are as follows:

1. The name consists of two words
2. The name of the element that is lower left in the Periodic Table is placed first
3. The second name has the ending *–ide*
4. To indicate the number of each atom, prefixes are used (except if there is only one atom of the first named-element when no prefix is used).

Thus to be able to name such compounds, it is necessary to know the names of the prefixes. These prefixes are listed in Table 7.4.

Table 7.4 The common prefixes used for naming binary covalent compounds.

number	prefix		number	prefix
1	mono-		6	hexa-
2	di-		7	hepta-
3	tri-		8	octa-
4	tetra-		9	nona-
5	penta-		10	deca-

For example, PCl_3 is called phosphorus *tri*chloride while PCl_5 is called phosphorus *penta*chloride. As another example, P_4O_{10}, is called *tetra*phosphorus *deca*oxide.

7.6 Differentiating Binary Ionic and Covalent Compounds

One of the challenges in naming binary compounds is to distinguish when to use ionic naming and when to use covalent naming. The change-over comes when the first-named element is a nonmetal or semimetal instead of a metal. The following Figure 7.3 shows some of the fluorides of the 3rd Period elements. As can be seen, the transition from metal to nonmetal occurs between aluminum and silicon.

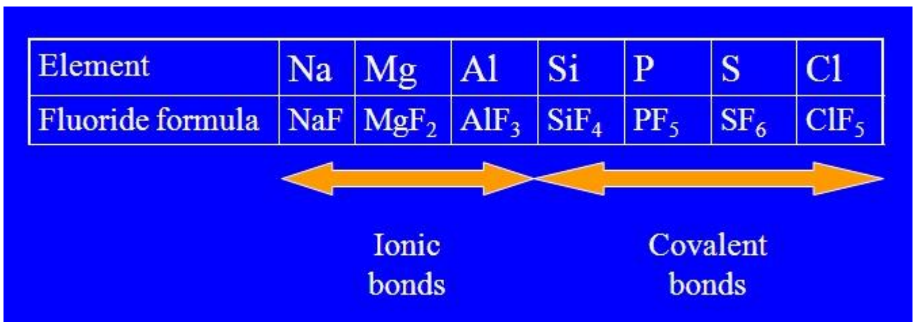


Figure 7.3 The bonding is ionic for the combination of metal and nonmetal but covalent for the combination of two nonmetals.

As the bonding changes, so does the nomenclature method. So for the metal-nonmetal combination, no prefixes are needed because there is only one possible formulation as sodium ion always has a charge of +1; magnesium ion, a charge of +2; and aluminum ion, a charge of +3. With the nonmetal-nonmetal combination, it is crucial to use prefixes. For example, in addition to the SF_6 shown, sulfur can form SF_4, SF_2, and S_2F_2. These four compounds would be named sulfur hexafluoride, sulfur tetrafluoride, sulfur difluoride, and disulfur difluoride. Figure 7.4 shows the pattern of names across the series shown in Figure 7.3.

Element	Na	Mg	Al	Si	P	S	Cl
Fluoride formula	NaF	MgF_2	AlF_3	SiF_4	PF_5	SF_6	ClF_5

Sodium fluoride

Magnesium fluoride

Aluminum fluoride

Silicon tetrafluoride

Phosphorus pentafluoride

Sulfur hexafluoride

Chlorine pentafluoride

Figure 7.4 The corresponding names for the Period 3 fluorides shown in Figure 7.3

Descending a group, the elements change from nonmetal to metal. Figure 7.5 shows the elements of Group 15, the change-over from nonmetal to metal occurring between arsenic (As) and antimony (Sb). Figure 7.5 also shows the different fluorides that they form. All but nitrogen form two different fluorides.

Element	Fluoride formulas
N	NF_3
P	PF_3, PF_5
As	AsF_3, AsF_5
Sb	SbF_3, SbF_5
Bi	BiF_3, BiF_5

Covalent bonds

Ionic bonds

Figure 7.5 Elements of Group 15, their fluorides, and the bonding transition from covalent to ionic.

As discussed in Section 7.4, for the nonmetals, prefixes are used for the number of each element. Using this method, the names of the fluorides of nitrogen, phosphorus, and arsenic can be deduced. The metals antimony and bismuth also form compounds with two different formulas. However, the naming of these, as described in Section 7.3, requires the use of Roman numerals to differentiate them. Antimony can be used as an example. For the compound, SbF_3, as the fluoride ion has a charge of −1, each antimony must have a charge of +3. Thus the compound would be named antimony(III) fluoride. For the compound, SbF_5, as the fluoride ion has a charge of −1, each antimony must have a charge of +5. Thus, this compound would be named antimony(V) fluoride. Figure 7.6 shows the names for each of the Group 15 fluorides.

Element	Fluoride formulas	
N	NF_3	Nitrogen trifluoride
P	PF_3, PF_5	Phosphorus trifluoride phosphorus pentafluoride
As	AsF_3, AsF_5	Arsenic trifluoride arsenic pentafluoride
Sb	SbF_3, SbF_5	Antimony(III) fluoride antimony(V) fluoride
Bi	BiF_3, BiF_5	Bismuth(III) fluoride bismuth(V) fluoride

Figure 7.6 The names of the fluorides of the Group 15 elements.

To summarise, naming binary compounds has to be undertaken with care. To illustrate, the examples of CaO, CuO, and CO will be used.

CaO – calcium is a metal. As it is combined with a nonmetal, the bonding will be ionic. Calcium has only one possible charge (+2), thus the compound is named calcium oxide.

CuO – copper is a metal. As it is combined with a nonmetal, the bonding will be ionic. However, copper can form ions of +1 or +2 charge. Oxygen forms an ion of charge −2, so each copper ion must have a charge of +2. This compound will be named copper(II) oxide.

CO – carbon is a nonmetal. As it is combined with a nonmetal, the bonding will be covalent. Thus prefixes need to be used. So this compound will be named carbon monoxide.

EXAMPLE 7.2

Determine the name of (a) SCl_2; (b) $SrCl_2$; and (c) $SnCl_2$.

Answer

(a) SCl_2 is a combination of two non-metals. For this covalent compound, the prefix system is used. So the name is sulphur dichloride.

(b) $SrCl_2$ is a combination of a metal and a non-metal. Strontium always has a charge of +2, so the name of this ionic compound is strontium chloride.

(c) $SnCl_2$ is also a combination of a metal and a non-metal. However, tin can form ions of charge +2 or +4. As the chloride ion always has a charge of –1, and there are two chloride ions, then the tin must have a charge of +2. The name of this ionic compound is tin(II) chloride.

7.7 The Formulas and Names of Polyatomic Ions

Up to now, binary neutral compounds have been discussed. First, it was binary ionic compounds, then it was binary covalent compounds. The only covalent compounds covered in this course are binary compounds. However, for ionic compounds, it is necessary to learn to name combinations of three and even four elements. These compounds contain polyatomic ions, that is, clusters of atoms held together by covalent bonds, but with a charge on the whole unit. In this Section, the formulas and corresponding names of some common polyatomic ions are listed, then in the following Section, the combining of cations and anions to form neutral compounds will be covered.

In Chapter 6, Section 6.6, the construction of electron-dot structures of a few polyatomic ions was shown. There are, in fact, many polyatomic ions and the structure of each one can be obtained using electron-dot symbols in the same way. To save time, it is much easier to memorize the formula and name of each of the ions, then in the following section, they will be matched with an ion of opposite charge to give a wider range of ionic compounds.

By far the most common type of polyatomic ion are the oxoanions. These are combinations of an element with oxygen atoms having an overall negative charge. Table 7.5 lists the common oxoanions by Group. The *–ate* ending was part of the naming system devised by de Morveau.

Table 7.5 Formulas and Names of Common Oxoanions

	Group 6	Group 7		Group 14	Group 15	Group 16	Group 17
Period 2				CO_3^{2-} carbonate	NO_3^- nitrate		
Period 3				SiO_4^{4-} silicate	PO_4^{3-} phosphate	SO_4^{2-} sulfate	ClO_4^- perchlorate
Period 4	CrO_4^{2-} chromate	MnO_4^- permanganate					

There are patterns to the formulas of the common oxoanions. For example, from left to right the negative charge on the oxoanion decreases by one. This trend is caused by the non-oxygen atom having one more electron across the sequence. Also, all of the oxoanions have four oxygen atoms except for those of the second Period which have three oxygen atoms. Finally, there are parallels in formula and name between the oxoanions of Group 6 and Group 16, and between those of Group 7 and Group 17.

For some elements, it is possible to have more than one oxoanion, each with a different number of oxygens. To indicate the one less oxygen atom, the ending is changed to *–ite*. There are two common examples shown in Table 7.6.

Table 7.6 Formulas and Names of Two Common Pairs of Oxoanions

Group 15	Group 16
NO_3^- nitrate	SO_4^{2-} sulfate
NO_2^- nitrite	SO_3^{2-} sulfite

For chlorine, there are four oxoanions known. To distinguish them, in addition to suffixes of –*ate* and –*ite*, prefixes of *per*– (for even more) and *hypo*– (for even less) are used. These four oxoanions are shown in Table 7.7.

Table 7.7 Formulas and Names of the Oxoanions of Chlorine

Group 17
ClO_4^- perchlorate
ClO_3^- chlorate
ClO_2^- chlorite
ClO^- hypochlorite

There are three other polyatomic ions that it is necessary to know, two of which were introduced in Chapter 6, Section 6.6. These are shown in Table 7.8.

Table 7.8 Formulas and Names of Other Polyatomic Ions

NH_4^+ ammonium
CN^- cyanide
OH^- hydroxide

7.8 The Formulas and Names of Ionic Compounds Containing Polyatomic Ions

In Section 7.4, it was shown how monatomic cations and monatomic anions combine in such a ratio as will form an electrically-neutral compound. The same procedure is applied in this Section to ionic compounds containing polyatomic ions.

COMPOUNDS OF OXOANIONS

As an example, to derive the formula for the ionic compound formed between aluminum ion and nitrate ion, the first step is to write the formulas of each ion: Al^{3+} and NO_3^-. To balance the charges, three NO_3^- ions will be needed. To clarify that it is three of the whole oxoanion, parentheses are used to give the formula of: $Al(NO_3)_3$. The name of this compound will be aluminum nitrate.

To write the name corresponding to a formula, it is crucial to check whether the cation has more than one possible charge (as was the case for binary ionic compounds in Section 7.2). For example, to deduce the name of the compound $Fe(OH)_2$, it is necessary to recall that iron can have a charge of +2 or +3. Each hydroxide ion has a charge of −1, thus to balance the total of 2(−1), that is, −2, the cation must have a charge of +2. The correct name would therefore be iron(II) hydroxide.

Parentheses are only used for formulas when there is more than one of the polyatomic ion. When the ratio of cation to anion is 1:1, it is important to be able to recognize the constituent ions in the formula. To illustrate, though the formula for ammonium perchlorate is written as NH_4ClO_4, there are the two constituent polyatomic ions in the compound: NH_4^+ and ClO_4^-.

ACID OXOANIONS

In Section 7.6, the formulas and names of common oxoanions was given. For oxoanions with more than one negative charge, it is possible to combine that anion with a positive hydrogen ion to give what is called an acid oxoanion. For example, sulfate ion, SO_4^{2-}, can combine with a hydrogen ion, H^+, to give the hydrogen sulfate ion, HSO_4^-. The formulas and corresponding names of common acid oxoanions are shown in Table 7.9. As there are two acid oxoanions of phosphorus, prefixes have to be used to distinguish them.

Table 7.9 Formulas and corresponding names of common acid oxoanions

Formula	Corresponding name
HCO_3^-	hydrogen carbonate
HSO_4^-	hydrogen sulfate
HSO_3^-	hydrogen sulfite
HPO_4^{2-}	monohydrogen phosphate
$H_2PO_4^-$	dihydrogen phosphate

COMPOUNDS OF ACID OXOANIONS

Such compounds are named by the regular rules for ionic compounds. The key to naming is to recognize the polyatomic acid-oxoanion. For example, to name $NaHCO_3$, it is crucial to realize that there are two ions in the compound, that is: Na^+ and HCO_3^-. The naming then becomes relatively easy: sodium hydrogen carbonate (this compound has a non-systematic name – confusingly – sodium bicarbonate).

EXAMPLE 7.3

Name each of the following compounds: (a) Na_2SO_4; (b) Na_2SO_3; (c) $NaHSO_4$; and (d) $FeSO_4$.

Answer

(a) Na_2SO_4 is a combination of the sodium cation (Na^+) and the sulfate anion (SO_4^{2-}). The name is sodium sulfate.

(b) Na_2SO_3 is a combination of the sodium cation (Na^+) and the sulfur oxoanion with one less oxygen atom – the sulfite anion (SO_3^{2-}). The name is sodium sulfite.

(c) $NaHSO_4$ is a combination of the sodium cation (Na^+) and the acid anion, hydrogen sulfate (HSO_4^-). The name is sodium hydrogen sulfate.

(d) The iron ion can have a charge +2 or +3. The sulfate ion (SO_4^{2-}) has a charge of 2–, so to balance the charge as they are in a 1:1 ratio, the iron cation in $FeSO_4$ must have a charge of +2. The name is iron(II) sulfate.

7.9 The Formulas and Names of Ionic Compounds Containing Water Molecules

Some ionic compounds have water molecules associated with them. These are called hydrated salts and can be defined as: solid ionic compounds in which molecules of water are trapped in the crystal lattice. For example, sodium carbonate is found in natural mineral deposits not as anhydrous (without water) sodium carbonate, but with ten molecules of water to each combination of two sodium ions and one carbonate ion. The formula is written as: $Na_2CO_3{\cdot}10H_2O$. To name hydrates, the word ‘hydrate’ is added after the name, with the number of water molecules identified by the same prefixes as used for binary covalent compounds (Table 7.10).

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Table 7.10 The common prefixes used for naming binary covalent compounds.

number	prefix		number	prefix
1	mono-		6	hexa-
2	di-		7	hepta-
3	tri-		8	octa-
4	tetra-		9	nona-
5	penta-		10	deca-

For example, $CaCl_2{\cdot}2H_2O$ is called calcium chloride *di*hydrate while $CuSO_4{\cdot}5H_2O$ is called copper(II) sulfate *penta*hydrate.

7.10 The Formulas and Names of Inorganic Acids

In Chapter 2, Section 2.6, one of the categories of compounds was acids. In this section, the naming of acids will be covered.

Acids are formed when certain hydrogen-containing compounds are dissolved in water. It is only when they are dissolved in water that they behave as acids and are named by the rules of acid-naming. For example, HCl is a gas which is named hydrogen chloride. When hydrogen chloride is dissolved in water, it forms an acid, HCl(*aq*). This acid is named hydrochloric acid. There are several binary acids, and their names are all constructed by putting the prefix *hydro-* and the suffix *-ic acid* around the stem name of the non-hydrogen element. Though the rules are for binary acids, the cyanide ion, CN^-, as it closely resembles the ions such as chloride, is named as if it was a monatomic ion. The common binary acids are listed in Table 7.11.

Table 7.11 Formulas and names of common binary acids.

Name	Corresponding formula
HF(*aq*)	hydrofluoric acid
HCl(*aq*)	hydrochloric acid
HBr(*aq*)	hydrobromic acid
HI(*aq*)	hydroiodic acid
HCN(*aq*)	hydrocyanic acid

Most of the common acids are derived from oxoanions. In Section 7.6, the formulas and names of oxoanions was covered. Then in Section 7.8, it was described how acid anions could be formed by adding one hydrogen ion to an oxoanion that had more than one negative charge. To form an acid from an oxoanion, the appropriate number of hydrogen ions is added to an oxoanion to balance the negative charge on the anion. For example, adding one hydrogen ion, H^+, to a nitrate ion, NO_3^-, gives the acid HNO_3(*aq*). Similarly, adding two hydrogen ions to a sulfate ion, SO_4^{2-}, gives the acid, H_2SO_4(*aq*). If there is only one oxoanion for that element then the name is formed by adding the ending *-ic acid*, just as the oxoanion itself had the ending *-ate*. The two common examples of these are shown in Table 7.12.

Table 7.12 A comparison of the formulas and names of the single oxoanions and the corresponding acids

Oxoanion			Corresponding acid	
Formula	Name		Formula	Name
CO_3^{2-}	carbonate		H_2CO_3(*aq*)	carbonic acid
PO_4^{3-}	phosphate		H_3PO_3(*aq*)	phosphoric acid

If there is a second oxoanion with one less oxygen atom then the ending is changed from *-ite* of the oxoanion to *-ous acid* of the corresponding acid. Table 7.13 shows the nitrogen oxoanions and the corresponding acids, while Table 7.14 shows the sulfur oxoanions and the corresponding acids.

Table 7.13 A comparison of the formulas and names of the nitrogen oxoanions and the corresponding acids

Oxoanion			Corresponding acid	
Formula	Name		Formula	Name
NO_3^-	nitrate		$HNO_3(aq)$	nitric acid
NO_2^-	nitrite		$HNO_2(aq)$	nitrous acid

Table 7.14 A comparison of the formulas and names of the sulfur oxoanions and the corresponding acids

Oxoanion			Corresponding acid	
Formula	Name		Formula	Name
SO_4^{2-}	sulfate		$H_2SO_4(aq)$	sulfuric acid
SO_3^{2-}	sulfite		$H_2SO_3(aq)$	sulfurous acid

In the case of the chlorine oxoacids, there are four, corresponding to the four oxoanions described in Section 7.6. The correlation is shown in Table 7.15.

Table 7.15 A comparison of the formulas and names of the chlorine oxoanions and the corresponding acids

Oxoanion			Corresponding acid	
Formula	Name		Formula	Name
ClO_4^-	perchlorate		$HClO_4(aq)$	perchloric acid
ClO_3^-	chlorate		$HClO_3(aq)$	chloric acid
ClO_2^-	chlorite		$HClO_2(aq)$	chlorous acid
ClO^-	hypochlorite		$HClO(aq)$	hypochlorous acid

EXAMPLE 7.4

Name each of the following compounds: (a) $HCl(g)$; (b) $HCl(aq)$; and (c) $HClO_4(aq)$.

Answer

(a) $HCl(g)$ – this is simply a binary covalent compound. The name is hydrogen chloride.

(b) $HCl(aq)$ – this is one of the common binary acids. The name is hydrochloric acid.

(c) $HClO_4(aq)$ – this is the chlorine oxoacid with the most oxygen atoms. The name is perchloric acid.

7.11 Where Next?

In Chapter 4, Section 4.4, the counting concept of the mole was introduced in the context of elements. Having now studied the composition of compounds, the next chapter deals with mole calculations involving compounds.