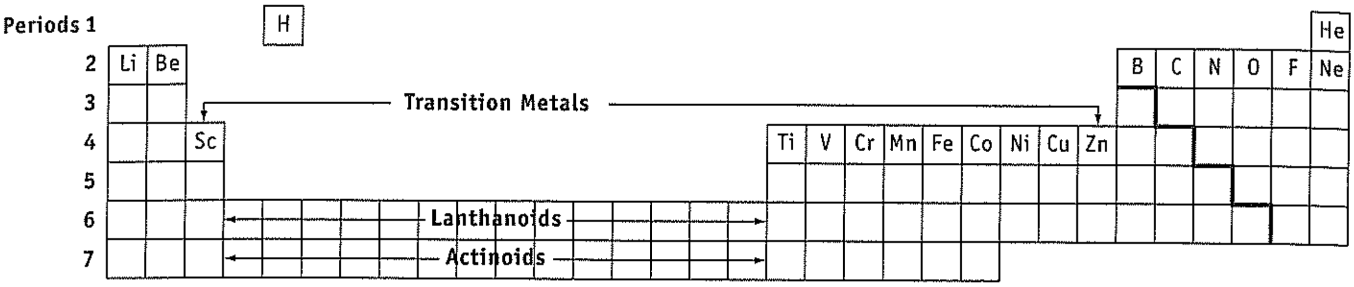
# Elements

# 2 The Periodic Table and organisation of elements

## Full Periodic Table

The layout of the full Periodic Table is shown below.




## Normal Periodic Table

The Periodic Table is usually in the shorter form shown on page 17.

### Periods

- *Periods* are the rows of the Periodic Table. The elements are arranged in order of increasing atomic number (number of protons in each atom's nucleus) across each Period. This is also roughly the order of increasing atomic mass.

**1** Complete the following table:

| Period | Number of elements in the Period | First element | Last element |
|---|---|---|---|
| 1 | 2 | | |
| 2 | | | Neon |
| 3 | | | |
| 4 | | | |
| 5 | | Rubidium | |
| 6 | | | |
| 7 | 23 | | Meitnerium |

### Groups

- *Groups* are the eight main columns. The Groups are usually numbered I–VIII using Roman numerals. The elements in each group have similar properties.
- Group I (alkali metals) begins with the metal lithium. Hydrogen has some chemical properties similar to these metals but because it is a gas it is shown separately.
- Group VII (halogens) comprises very chemically reactive non-metals.
- Group VIII (noble gases) contains gases which are inert (unreactive). These gases are the only elements in which the particles are discrete atoms (separate atoms).

**2** True or false?

(a) Sodium and sulfur have similar properties.
(b) Fluorine is chemically similar to chlorine.
(c) A helium balloon contains helium molecules.
(d) Light globes are filled with argon because it does not react with the hot metal filament.
(e) Calcium reacts the same way as magnesium does in many chemical reactions.

- **The zig-zag line in the Periodic Table separates metals from non-metals.** Metals are on the left of the line and non-metals are on the right. Elements close to the line are semi-metals and have both metal and non-metal characteristics.

**3** Classify the following elements as metals, semi-metals, or non-metals:

| | | | | | |
|---|---|---|---|---|---|
| argon | arsenic | boron | bromine | germanium | gold |
| silicon | silver | sodium | sulfur | tellurium | titanium |

- **Transition metals make up the block in the middle of the table.** Many transition metals are used in industry. Compounds of transition metals are often coloured.

# Isotopes and relative atomic mass

## Mass number and atomic number

- The *mass number (A)* of an atom is the total number of protons and neutrons contained in the nucleus of the atom.
- The *atomic number (Z)* is the number of protons only.

mass number / atomic number ELEMENT e.g. $^{23}_{11}Na$

Number of neutrons in an atom = Mass number − Atomic number

**1** Complete the following table:

| Element atom | Symbol | No. of protons | No. of neutrons |
|---|---|---|---|
| Helium | $^{4}_{2}He$ | | |
| Sodium | $^{23}_{11}Na$ | | |
| Oxygen | $^{16}_{8}O$ | | |
| Phosphorus | $^{31}_{15}P$ | | |
| Copper | $^{63}_{29}Cu$ | | |
| Iodine | $^{127}_{53}I$ | | |

- *Isotopes* are different atoms of the same element which have different numbers of neutrons but the same number of protons in the nucleus.
- Many elements have naturally occurring isotopes. Isotopes can be made artificially by bombarding atoms of the element with neutrons.
- Examples of isotopes:

1. Isotopes of hydrogen:
   - $^{1}_{1}H$ — pronounced 'Hydrogen 1'
   - $^{2}_{1}H$ — Hydrogen 2 (deuterium)
   - $^{3}_{1}H$ — Hydrogen 3 (tritium)
2. Isotopes of carbon:
   - $^{12}_{6}C$ — Carbon 12
   - $^{13}_{6}C$ — Carbon 13
   - $^{14}_{6}C$ — Carbon 14

**2** (a) What do all three isotopes of hydrogen have in common? ______________________________

(b) What difference is there in the atomic structure of hydrogen, deuterium and tritium?

______________________________________________________________

(c) How many protons are there in the nucleus of every carbon atom? ____________

(d) The most common isotope of carbon has 6 neutrons. What is its mass number? ____________

## Atomic mass

- Almost all the mass of an atom is in its nucleus. A neutron has very nearly the same mass as the mass of a proton. Electrons have almost zero mass.
- The real mass of a proton is $1.673 \times 10^{-24}$ g and of an electron, $9.110 \times 10^{-28}$ g. Atoms are too small to be weighed easily.

## Relative atomic mass

- *Relative atomic mass* ($A_r$) is the mass of an atom compared with a standard atom. The relative atomic mass value of sodium is 23: a sodium atom is 23 times heavier than the lightest hydrogen atom ($^{1}_{1}H$) which has a relative atomic mass value of 1. The modern standard atom is the $^{12}_{6}C$ isotope which has a value of 12.

**3** Given that a proton and a neutron each have a mass value of 1, what is the relative atomic mass of:

(a) the $^{3}_{1}H$ isotope? ____________

(b) an atom containing 14 neutrons and 13 protons? ____________

(c) an atom which has mass number 9? ____________

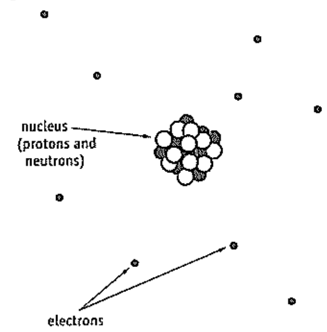
(d) the deuterium isotope? ____________

(e) $^{34}_{16}S$ ? ____________

# 3 Atomic structure
# Atoms and ions

## Atoms

- The structure showing main subatomic particles is shown in the diagram below:




- The central nucleus is only about 1/10 000 of the atom's diameter.
- The nucleus contains:
  *protons* — positive particles (symbol p+)
  *neutrons* — neutral particles (symbol n)
- The space around the nucleus is occupied by fast-moving **electrons** — negative charges (symbol $e^-$).
- In any complete atom the number of positive protons is balanced by an equal number of negative electrons. A complete atom is electrically neutral.
- The *atomic number* (*Z*) of an element is the *number of protons* in its nucleus. For example, a hydrogen atom has 1 proton, a chlorine atom has 17 protons and a uranium atom has 92 protons.

**1** Use the Periodic Table on page 17 to complete the following table:

| Element atom | | | Copper | | Chlorine |
|---|---|---|---|---|---|
| Symbol | | Ca | | | |
| No. of $p^+$ | 19 | | | | |
| No. of $e^-$ | | | | 8 | |

## Ions

- *Ions* are charged particles. When atoms gain or lose electrons they change into ions. Extra electrons change a neutral atom into a negative ion. Losing electrons changes a neutral atom into a positive ion. The number of protons in the nucleus does not change when ions form.
- **Positive ions (cations) form when metal atoms lose electrons.** A complete atom has an equal number of protons and electrons. The positive charge is caused by the number of protons exceeding the number of electrons remaining. The name of a metal ion is the same as the name of the metal atom.
- **Negative ions (anions) form when non-metal atoms gain electrons.** The negative charge is caused by the number of electrons exceeding the number of protons. The names of non-metal ions end in 'ide', e.g. chloride ion.
- **General guide to the number of electrons lost or gained by atoms which form ions:**
  1. Group I metals lose 1 $e^-$; ions have a single positive charge, e.g. $K^+$.
  2. Group II metals lose 2 $e^-$; ions have two positive charges, e.g $Ca^{2+}$.
  3. Group III metals lose 3 $e^-$; ions have three positive charges, e.g. $Al^{3+}$.
  4. Group VI non-metals gain 2 $e^-$; ions have two negative charges, e.g. $O^{2-}$.
  5. Group VII non-metals gain 1 $e^-$; ions have a single negative charge, e.g. $Cl^-$.

  (There is no general rule for the charge on ions formed from atoms of Groups IV and V. Atoms of Group VIII elements do not form ions or combine with other atoms in any way.)

**2** Complete the table.

| Ion | No. of $p^+$ | No. of $e^-$ |
|---|---|---|
| $Na^+$ | | |
| $Ca^{2+}$ | | |
| $Al^{3+}$ | | |
| $F^-$ | | |
| $O^{2-}$ | | |
| $Zn^{2+}$ | | |




# 3 Atomic structure
# Atoms and ions (cont.)

**3** Complete the following tables of ion names and symbols:

| Metal ion (cation) | Sodium | | Caesium | | Calcium | Strontium | | |
|---|---|---|---|---|---|---|---|---|
| Symbol | | $K^+$ | | $Mg^{2+}$ | | | $Ba^{2+}$ | $Al^{3+}$ |

| Non-metal ion (anion) | Oxide | | | Chloride | | |
|---|---|---|---|---|---|---|
| Symbol | | $S^{2-}$ | $F^-$ | $Cl^-$ | $Br^-$ | $I^-$ |

## Other common metal ions

- As well as the metals in Groups I, II and III, many common metals are located in the transition metal block of the Periodic Table. Tin and lead, at the bottom of Group IV, are also widely used. Common metal ions not included in the table above are listed below. Some metals can form more than one ion. Roman numerals are included in the names of these ions to indicate the number of positive charges on the ion.

**4** Complete the names missing from the table below:

| Ion name | Symbol | Ion name | Symbol |
|---|---|---|---|
| ______ | $Co^{2+}$ | Iron (III) | $Fe^{3+}$ |
| Cobalt (III) | $Co^{3+}$ | ______ | $Pb^{2+}$ |
| ______ | $Cu^+$ | Tin (II) | $Sn^{2+}$ |
| Copper (II) | $Cu^{2+}$ | Mercury (II) | $Hg^{2+}$ |
| Gold (I) | $Au^+$ | Nickel | $Ni^{2+}$ |
| ______ | $Au^{3+}$ | Silver | $Ag^+$ |
| ______ | $Fe^{2+}$ | Zinc | $Zn^{2+}$ |

## Ionic bonds

- In any chemical reaction, bonds are formed between atoms. Only the electrons of the atoms are involved in joining the atoms; the nucleus is unaffected.
- Generally, when metals combine with non-metals, each non-metal atom pulls a particular number of electrons from the metal atom. This creates negative and positive ions.
- Once the positive and negative ions have formed, they are held together by an ionic bond — the strong electrostatic attraction between ions with opposite charges.

**5** Complete the table below. Name the *atoms* which have combined by ionic bonding to form each of the compounds. Name and write symbols for the positive and negative *ions* in the compound.

| Atoms combining | Ionic compound | Positive ion (name, symbol) | Negative ion (name, symbol) |
|---|---|---|---|
| | Sodium bromide | | |
| | Lead (II) iodide | | |
| | Potassium sulfide | | |
| | Calcium fluoride | | |
| | Barium chloride | | |
| | Gold (I) oxide | | |
| | Iron (III) oxide | | |

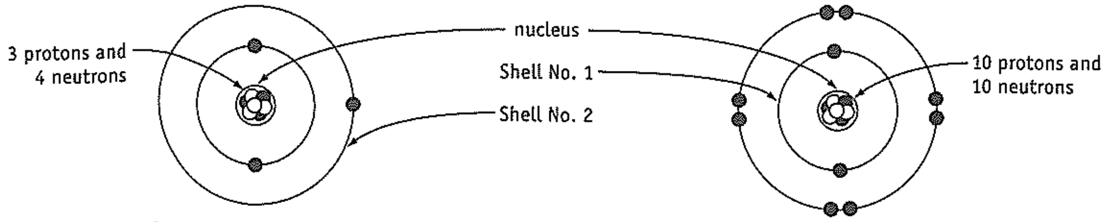
Periodic Table patterns

# 1 Electron arrangement and atomic radius

## Electron shells

- The electrons moving in the space around the nucleus of an atom occupy different energy levels. This idea is simplified in diagrams which show the electrons in orbits around the nucleus. The energy levels are called shells and are numbered from the nucleus outwards. Shell No. 1 is the lowest energy level.

**Example: Lithium atom** **Example: Neon atom**




- Each complete lithium atom has 3 electrons.
- As shown in the orbit-type diagram, the electron arrangement is:
  1. Shell No. 1: 2 $e^-$
  2. Shell No. 2: 1 $e^-$
- The electron configuration of lithium is 2, 1.

- Every neon atom has 10 electrons.
- The electron arrangement is:
  1. Shell No. 1: 2 $e^-$
  2. Shell No. 2: 8 $e^-$
- The electron configuration of neon is 2, 8.

**1** On a separate sheet of paper, draw orbit-type diagrams of the electron arrangement in atoms of elements between lithium and neon. (Across Period 2, each successive electron goes into the second shell.)

## Number of electrons per shell

- The shells can be compared to shelves in a bookcase. The shelves can be empty or they can have books on them. The books cannot be between shelves. There is a maximum number of books which can fit on each shelf. Similarly, electrons cannot be between shells and there is a maximum number of electrons which can fit in each shell.

$$\text{Maximum number of electrons} = 2n^2$$

  where $n$ = the shell number (Shell No. 1, Shell No. 2, etc.)
- When the number of electrons in the third shell reaches 8 (argon atoms), electrons start to fill the fourth shell.
- The electron configuration of potassium atoms is 2, 8, 8, 1. There are places left for more electrons in the third shell. These remaining third shell vacancies are filled by electrons of transition metal atoms.

**2** (a) Compile a table using the following headings for elements in order from atomic no. 1 to no. 20.

| Element | Atomic no. | 1st shell $e^-$ | 2nd shell $e^-$ | 3rd shell $e^-$ | 4th shell $e^-$ |
|---|---|---|---|---|---|
| Hydrogen | 1 | 1 | | | |

(b) After argon, how many places remain to be filled in the third shell? ______________

## Atomic radius

- The nucleus is only about 1/10 000 of an atom's radius. The size of an atom is determined by the space taken up by its electrons. The more electron shells, the bigger the atom.
- The Period number indicates how many electron shells are occupied in atoms in that period, e.g. atoms in Period 1 have 1 shell occupied; atoms in Period 7 have 7 shells occupied.
- **Trend: Atomic radius increases down each Group of the Periodic Table.**

**3** Refer to electron configuration to explain why:

(a) the smallest atom in Group VII is fluorine and the largest is astatine; ______________

______________

(b) silicon atoms have larger diameters than nitrogen atoms. ______________

______________

# Periodic Table patterns

# 3 Electron configuration and valency

## Electron configuration

- In orbit-type diagrams of atomic structure, electrons are shown in layers. When atoms join together, electrons in the outer layer interact. The outer shell is called the valence shell. Electrons in the outer shell are the valence electrons.
- Electron configuration of Group VIII elements:
  1. He...2
  2. Ne...2, 8
  3. Ar...2, 8, 8
  4. Kr...2, 8, 18, 8
  5. Xe...2, 8, 18, 18, 8
  6. Rn...2, 8, 18, 32, 18, 8
- Group VIII elements (Noble Gases):
  1. have the highest ionisation energy in each row of the Periodic Table; it is very difficult to remove an electron from the outer shell of their atoms;
  2. are inert (unreactive); their outer shell electrons do not interact with valence electrons of other atoms;
  3. are monatomic gases composed of individual atoms; their atoms do not join together to make molecules.

**1** (a) Which Noble Gases have complete outer shells? ______________________

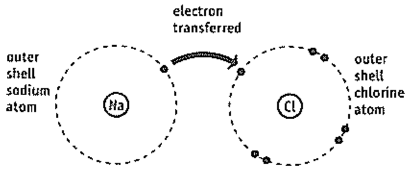
(Formula: Maximum number of electrons = $2n^2$ where $n$ is the shell number.)

(b) Select words from the word bank and fill in the spaces in the sentences below:

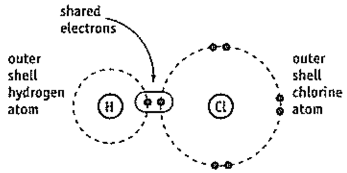
| | | | | | |
|---|---|---|---|---|---|
| although | neon | argon | krypton | xenon | each |
| lack | incomplete | eight | compounds | feature | case |

The Noble Gases do not form chemical ____________; they are chemically inert. Their atoms do not even join to ____________ other. Their ____________ of reactivity indicates that the outer shells of all Noble Gas atoms are complete but this is not the ____________. ____________ the outer shells of helium and ____________ contain the maximum number of electrons, those of ____________, ____________, ____________ and radon are incomplete. All those with ____________ outer shells have eight electrons in the outer shell. This common ____________ of these unreactive elements suggests that ____________ is a stable number of electrons.

- **Atoms combine to achieve a stable number of electrons in the outer shell.**
  Stable number of electrons in the outer shell = 8.
  For small atoms with only one electron shell, stable number = 2.
- Two ways in which atoms achieve a stable number of outer shell electrons are as follows:
  1. by transferring valence electrons. Metals combine with non-metals in this way.
  2. by sharing valence electrons with other atoms. Combinations of non-metals occur in this way.




**Ionic bond: strong attraction between positive and negative ions**




**Covalent bond: atoms held together by pair of electrons moving between them**






# 3 Electron configuration and valency (cont.)

## Valency

- The *valency* of an element is the number of electrons which its atom needs to gain, lose or share with another atom to achieve a stable number in the outer shell. Noble Gases have stable numbers of electrons. Their valency is 0.
- Some examples are:
  1. Sodium (metal) has a valency of 1. After a sodium atom loses 1 electron to a non-metal atom it has a stable shell of 8 electrons.
  2. Chlorine (non-metal) has a valency of 1. By gaining 1 electron from a metal or by sharing 1 electron with another non-metal, a chlorine atom will have the stable number of 8 electrons in its outer shell.
- **Trends in valency:**

| Group | I | II | III | IV | V | VI | VII | VIII |
|---|---|---|---|---|---|---|---|---|
| Valency | 1 | 2 | 3 | 4 | 3 | 2 | 1 | 0 |

**2** *Metals and non-metals combining*

(a) How many valence electrons are there in:

(i) a magnesium atom? ____________ (ii) an oxygen atom? ____________

(b) When magnesium joins to oxygen, $Mg^{2+}$ ions and $O^{2-}$ ions form.

(i) How many electrons were transferred? ____________

(ii) Which atom lost the electrons? ____________

(iii) Which atom gained the electrons? ____________

(c) How many electrons are there in the outer shell of:

(i) a $Mg^{2+}$ ion? ____________ (ii) an $O^{2-}$ ion? ____________

**3** *Non-metals combining* Molecules of the compound ammonia form when nitrogen and hydrogen atoms combine by sharing electrons. The valency of hydrogen is 1.

(a) In which Group is nitrogen? ____________

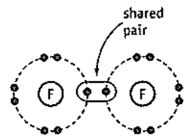
(b) What is the valency of nitrogen? ____________

(c) How many electrons does a nitrogen atom need to make a stable outer shell? ____________

(d) How many electrons does each hydrogen atom have available for sharing? ____________

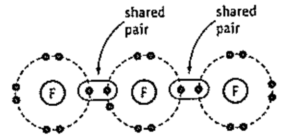
(e) How many hydrogen atoms share electrons with one nitrogen atom? ____________

**4** Diatomic elements: hydrogen, nitrogen, oxygen, fluorine, chlorine, bromine and iodine are composed of molecules in which each molecule contains two atoms. The two atoms are joined by a covalent bond. Diagram A shows how two fluorine atoms combine by sharing electrons. Diagram B shows how three fluorine atoms *might* join.




A

only outer shell electrons shown




B

(a) Including the shared electrons, how many electrons does each fluorine atom have? ____________

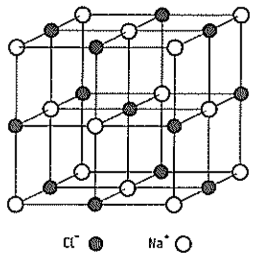
(b) Is this likely to happen? Explain your answer. ____________

# Ionic compounds and empirical formulas

## Ionic compounds

- These are composed of positive ions and negative ions in an ionic lattice which is a type of infinite array; there are no separate particles.
- The positive and negative ions are held together by ionic bonds.
- An ionic bond is the electrostatic attraction between a positive and a negative ion.
- A crystal of ionic compound is electrically neutral because the total number of positive charges equals the total number of negative charges.
- Ionic compounds usually form when metals combine with non-metals.

**Example: Sodium chloride composed of $Na^+$ and $Cl^-$ ions**




equal numbers of $Na^+$ ions and $Cl^-$ ions

**1** (a) What is the size of the charge on each positive sodium ion? ______________

(b) What is the size of the charge on each negative chloride ion? ______________

(c) Explain why there must be equal numbers of sodium and chloride ions in a sodium chloride crystal.

______________________________________________

**2** Explain why there are two $F^-$ ions for every one $Ca^{2+}$ in the ionic compound, calcium fluoride.

______________________________________________

______________________________________________

## Formulas of ionic compounds

- The chemical formula of an ionic compound shows the simplest ratio of the number of positive and negative ions.
  In a piece of sodium chloride there may be millions of $Na^+$ and $Cl^-$ ions but there are always equal numbers of $Na^+$ and $Cl^-$ ions. The simplest ratio is $Na^+$ to $Cl^-$ = 1:1.
- The chemical formula of sodium chloride is $Na_1Cl_1$.
- Usually the subscript '$_1$' is left out and the formula is written as NaCl. (Chemists agree that if there is no subscript written it means '$_1$'.)
- Empirical formulas are chemical formulas which show only the simplest ratio of atoms (or ions) and not the actual ratio.
  All formulas for ionic compounds are empirical formulas.

**3** (a) Write the chemical formula for calcium fluoride. ______________

(b) Why is this an empirical formula? ______________________________

**4** (a) What is the simplest ratio in which $Mg^{2+}$ and $O^{2-}$ could combine to form an ionic compound? ______________

(b) Write the formula for the compound. ______________






# Ionic compounds and empirical formulas (cont.)

## Radicals

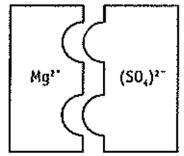
- Radicals are charged groups of atoms. Another name for them is 'polyatomic ions'. Some examples are:
  1. $(NO_2)^-$ nitrite ion
  2. $(CH_3COO)^-$ ethanoate ion
  3. $(H_2PO_4)^-$ dihydrogen phosphate ion
- Brackets are placed around the group of atom symbols to emphasise that the whole group has a charge. In some formulas for ionic compounds made from radicals, the brackets can be left out. However, it is best to always include the brackets until you are very familiar with writing chemical formulas.
- Some examples of ionic compounds containing radicals are:
  1. $Na(NO_2)$ sodium nitrite
  2. $Ca(H_2PO_4)_2$ calcium dihydrogen phosphate

## Naming ionic compounds

- Ionic compounds always have two-part names.
- The name of the positive ion comes first.
- The positive ion has the same name as its atom, e.g. $Na^+$ is sodium ion.
- The only common positive radical is ammon*ium* $(NH_4)^+$.
- The name of the negative ion comes second.
- The ending of the atom is changed to 'ide' to name simple negative ions, e.g $Cl^-$ is chlor*ide* ion.
- The names of negative radicals usually end in 'ate' or 'ite'; the exception $(OH)^-$ is hydroxide ion.

**5** You will need a minimum of three copies of page 46 to do this exercise.
Example question: What is the formula for the ionic compound, magnesium sulfate?
How to use the page 46 sheet:

- Cut out magnesium ions (positive) and sulfate ions (negative). Join sufficient magnesium and sulfate ions together to make the smallest rectangle possible (see the diagram at right).
- The number of each type of ion in this smallest possible rectangle gives you the empirical formula of magnesium sulfate...$Mg(SO_4)$.




(a) Use the symbols on page 46 to write chemical formulas for the following ionic compounds:

| | | |
|---|---|---|
| iron II sulfide ________ | zinc carbonate ________ | iron III sulfate ________ |
| ammonium iodide ________ | sodium carbonate ________ | potassium phosphate ________ |
| aluminium nitrate ________ | ammonium chloride ________ | magnesium hydroxide ________ |
| silver oxide ________ | copper II oxide ________ | copper II carbonate ________ |
| lead iodide ________ | sodium hydrogen carbonate ________ | calcium phosphate ________ |

(b) (i) After part (a) you will have cut-out symbols remaining. Join positive and negative ions together to make as many electrically neutral ionic compounds as you can.

(ii) List the names and empirical formulas of the compounds you have made:

| Name | Formula | Name | Formula |
|---|---|---|---|

Chemical formulas

# 2 Cut-outs for exercises in 'Ionic compounds and empirical formulas'

- $Al^{3+}$ aluminium
- $Fe^{3+}$ iron III
- $Ca^{2+}$ calcium
- $Mg^{2+}$ magnesium
- $Zn^{2+}$ zinc
- $Fe^{2+}$ iron II
- $Pb^{2+}$ lead
- $Cu^{2+}$ copper II
- $Na^{+}$ sodium
- $K^{+}$ potassium
- $Ag^{+}$ silver
- $(NH_4)^{+}$ ammonium

- nitrate $(NO_3)^{-}$
- hydrogen carbonate $(HCO_3)^{-}$
- fluoride $F^{-}$
- chloride $Cl^{-}$
- bromide $Br^{-}$
- iodide $I^{-}$
- sulfide $S^{2-}$
- hydroxide $(OH)^{-}$
- oxide $O^{2-}$
- carbonate $(CO_3)^{2-}$
- sulfate $(SO_4)^{2-}$
- phosphate $(PO_4)^{3-}$

# Bonding and physical properties

## Structural classification

- Chemical substances can be classified according to their bonding structure. Different structures have different patterns of physical properties.

| Bonding structure | Type of substance | Conduct electricity in... solid state | liquid state | aqueous solution | MP (°C) | BP (°C) |
|---|---|---|---|---|---|---|
| Metallic lattice | Elements, e.g. Fe | Yes | Yes | No, do not dissolve | Variable, e.g. Hg −39°, Na 98° Cu 1085°, W 3410° | Very high, generally >2000° Hg 357° Na 883° Cu 2572° W 5660° |
| Molecular (separate molecules) | Elements, e.g. $O_2$ | No | No | No | Low, most <0° $S_8$ 113° $P_4$ 44° | Low, most <60° $S_8$ 445° $P_4$ 280° |
| Monatomic (separate atoms) | Elements, e.g. Ne | No | No | No | Very low, < −100° | Very low, < −100° |
| Covalent network | Elements, e.g. Si | Graphite: yes Diamond: no Si, Ge: slightly | No | No, do not dissolve | High, Si 1410° Ge 937° diamond, graphite sublime >3550° | Very high, Si 3267° Ge 2830° |
| Ionic lattice | Compounds, e.g. NaCl | No | Yes | Yes | High 300°–1500° | Most decompose before boiling |
| Molecular (separate molecules) | Compounds, e.g. $H_2O$, $NH_3$, glucose | No | No | Variable, conduct if molecule dissociates, e.g. HCl | Moderate, generally <100° | Moderate, generally <300° |
| Covalent network | Compounds e.g. $SiO_2$ | No | No | No, do not dissolve | High, $SiO_2$ 1610° | Very high, $SiO_2$ 2230° |

### Comparing different structures of elements

- In both metallic lattice and covalent network, the structure extends throughout the sample; there are no separate particles. Also, both these structures typically have very high boiling points, indicating strong forces between the particles. However, metallic lattice structure conducts electricity in the solid and liquid states but covalent network structure does not conduct in either state. (The exception is graphite, a covalent network. Graphite conducts because delocalised *p* orbital electrons in the atoms are free to move.)
- Monatomic and molecular elements are similar in being non-conductors of electricity. Monatomic elements all have very low boiling points, but molecular elements have a wide range of boiling points.

**1** (a) Highlight the words in the paragraphs above which are used to make comparisons between different types of bonding ('similar' is one example).

(b) Use the same words to write sentences comparing the different structures of compounds.

Chemical equations

# 1 Writing and balancing equations

## Reactants → products

- The reactants are the chemical substances present before a chemical reaction occurs.
- The products are the new chemical substances formed by the reaction.
- In a chemical equation the change of reactants into products is shown by an arrow.
- The reactants and products of a chemical reaction are pure substances (elements or compounds). They are shown in the equation by their chemical formulas. For example, the chemical equation for the reaction between magnesium metal and dilute sulfuric acid to produce magnesium sulfate and hydrogen is:

$$Mg + H_2(SO_4) \rightarrow Mg(SO_4) + H_2$$

*Note:* Students who are very familiar with writing formulas may omit the brackets around the radical in those cases where only one radical group is present in the compound formula.

## Subscripts

- These are used in equations to indicate the state (solid, liquid, gas or dissolved in water) of the reactants and products. For example:

$$Mg_{(s)} + H_2(SO_4)_{(aq)} \rightarrow Mg(SO_4)_{(aq)} + H_{2(g)}$$

This equation indicates that the sulfuric acid and magnesium sulfate are aqueous (dissolved in water), the hydrogen is a gas and the magnesium is solid.

**1** Read the following passage and write a chemical equation, including subscripts, for the reaction described. (The formulas for hypochlorous acid and sodium hypochlorite are HClO and NaClO, respectively.)

*Bleach, a household disinfectant, contains sodium hypochlorite. The sodium hypochlorite reacts with water to produce hypochlorous acid and sodium hydroxide, which both remain dissolved in the water. The hypochlorous acid kills any micro-organisms which might be in the water.*

___

___

## Types of reactions

### Decomposition reactions

- This type of reaction has only one compound as the reactant. During the reaction the atoms in the compound are rearranged to form two or more products. The products can be elements or compounds. For example:

$$Cu(CO_3)_{(s)} \rightarrow CuO_{(s)} + CO_{2(g)}$$

**2** (a) Name the reactant and products and give their states as indicated in the above decomposition reaction.

___

(b) Why must the single reactant in a decomposition reaction always be a compound?

___

### Synthesis reactions

- In this type of reaction only one product is formed from two or more reactants. The product is 'synthesised' or built up from the reactants. For example:

$$Fe_{(s)} + S_{(s)} \rightarrow FeS_{(s)}$$

*Note:* The formulas for all metals and for most solid non-metals are the symbols for these elements.

# Writing and balancing equations (cont.)

### Combustion reactions

Combustion is another name for burning. This type of reaction is characterised by the release of large amounts of heat and light when an element or a compound reacts with oxygen. Some examples are:

1 If it is heated first, sulfur burns in air with a violet-coloured flame. The sulfur combines with oxygen in the air. The flame is caused by the release of large amounts of heat and light energy.

$$S_{(s)} + O_{2(g)} \rightarrow SO_{2(g)} \quad \text{heat and light energy released}$$

2 White phosphorus burns spontaneously in air; it does not have to be heated to undergo a combustion reaction with the oxygen contained in air.

$$P_{4(s)} + 5\,O_{2(g)} \rightarrow P_4O_{10(s)} \quad \text{heat and light energy released}$$

Notice that in the equation for the combustion of white phosphorus, the number 5 is before the formula for oxygen. This number is necessary to balance the equation. The balanced equation shows that 1 phosphorus molecule reacts with 5 oxygen molecules to produce 1 tetraphosphorus decaoxide ($P_4O_{10}$) molecule.

## Balancing equations

- Without the balancing number (5) the above equation is: $P_{4(s)} + O_{2(g)} \rightarrow P_4O_{10(s)}$
- On the reactant side (left side) of the equation: total number of phosphorus atoms = 4 and total number of oxygen atoms = 2.
- On the product side (right side) of the equation: total number of phosphorus atoms = 4 and total number of oxygen atoms = 10.
- The equation in this form is unbalanced. There are more oxygen atoms on the right side than there are on the left side of the equation. This is not possible; the oxygen atoms in the product molecules can only come from the reactant molecules. To balance any equation, numbers are placed before the formulas to make the total number of each type of atom equal on both sides of the equation.

### How to balance an equation

*Example:* $Mg_{(s)} + O_{2(g)} \rightarrow MgO_{(s)}$

Step 1: Count the total number of each type of atom on the left side of the equation: total Mg atoms = 1 and total O atoms = 2.

Step 2: Count the total number of each type of atom on the right side of the equation: total Mg atoms = 1 and total O atoms = 1.

Step 3: Add a balancing number before MgO to make the numbers of oxygen atoms equal on both sides: $Mg_{(s)} + O_{2(g)} \rightarrow 2\,MgO_{(s)}$

Step 4: Check the total number of each type again: Mg...1 atom on the left and 2 atoms on the right and O...2 atoms on each side.

Step 5: Add another balancing number to correct the imbalance in the numbers of magnesium atoms:

$$2\,Mg_{(s)} + O_{2(g)} \rightarrow 2\,MgO_{(s)}$$

Step 6: Check the total number of each type again: Mg...2 atoms on each side; O...2 atoms on each side. The equation is balanced!

*Another example:* $N_{2(g)} + H_{2(g)} \rightarrow NH_{3(g)}$

Steps 1, 2 and 3: $N_{2(g)} + H_{2(g)} \rightarrow 2\,NH_{3(g)}$

Steps 4 and 5: $N_{2(g)} + 3\,H_{2(g)} \rightarrow 2\,NH_{3(g)}$

Step 6: Check again — the equation is balanced!

**3** (a) Check each of the following equations. If the equation is unbalanced, add balancing numbers *before* the formulas.

(i) $Cu_{(s)} + O_{2(g)} \rightarrow CuO_{(s)}$

(ii) $Na_{(s)} + H_2O_{(l)} \rightarrow Na(OH)_{(aq)} + H_{2(g)}$

(iii) $PbO_{(s)} \rightarrow Pb_{(s)} + O_{2(g)}$

(iv) $H_2O_{2(l)} \rightarrow H_2O_{(l)} + O_{2(g)}$

(v) $CH_{4(g)} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(g)}$ heat and light released

(b) Identify combustion, decomposition and synthesis in the equations (i) to (v) above.

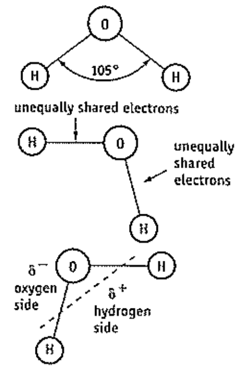





# Dissociation and precipitation

## Polar water molecules

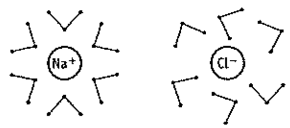
- Water in any state (solid, liquid or gas) is composed of molecules. In the $H_2O$ molecule each hydrogen atom is joined to the oxygen atom by a covalent bond. The angle between the covalent bonds is 105°.
- Oxygen has a greater electron affinity than hydrogen does. The shared pair of electrons in each covalent bond spend more time around the oxygen atom than they do around the hydrogen atom.
- Each covalent bond is polar; the oxygen end is slightly negative and the hydrogen end is slightly positive.
- Because the water molecule contains polar covalent bonds and has a bent shape, the oxygen side of a water molecule has a small negative charge and the hydrogen side has a small positive charge.




## Dissociation of ionic compounds in water

- When an ionic compound (e.g. NaCl) is added to water it usually dissolves. This occurs because the polar water molecules attract the ions in the sodium chloride. The attraction is strong enough to break the ionic bonds which hold the $Na^+$ and $Cl^-$ ions together in the solid ionic lattice structure. The $Na^+$ and $Cl^-$ ions dissociate (separate) when water is added to solid sodium chloride. Each $Na^+$ and each $Cl^-$ ion becomes hydrated (surrounded by water molecules).

Hydration of ions




Water molecules surround ions

**1** (a) Lead chloride is an example of an ionic compound which is insoluble in water. Compare the attractive force between $Pb^{2+}$ and $Cl^-$ ions with the attractive force between polar water molecules and these ions.

______________________________

(b) Name three other ionic compounds which are insoluble in water. ______________________________

### Net ionic dissociation equations

- The full equation for the dissociation of sodium chloride into ions in water is:

$$NaCl_{(s)} + x\,H_2O_{(l)} \rightarrow Na^+{}_{(aq)} + Cl^-{}_{(aq)} + x\,H_2O_{(l)}$$

where $x$ = an unknown number of water molecules

- The water molecules are not changed by the dissociation reaction although they do cause the ionic compound to dissociate. Usually the water molecules are left out and the ionic equation is written in its shorter ('net') form. The subscript $_{(aq)}$ indicates the involvement of water molecules.

$$NaCl_{(s)} \rightarrow Na^+{}_{(aq)} + Cl^-{}_{(aq)}$$

- Where the ratio of positive to negative ions in the ionic lattice is not 1:1 (e.g. $K_2(CO_3)$), the net ionic dissociation equation must show the correct ratio of these ions in solution (e.g. two potassium ions for every one carbonate ion).

$$K_2(CO_3)_{(s)} \rightarrow 2\,K^+{}_{(aq)} + (CO_3)^{2-}{}_{(aq)}$$

**2** Write net ionic dissociation equations for the dissolving in water of each of the following ionic compounds:

(a) potassium iodide ______________________________

(b) lead nitrate ______________________________

(c) calcium chloride ______________________________

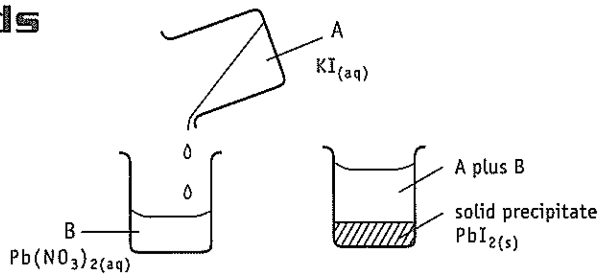
(d) ammonium hydroxide ______________________________

(e) sodium sulfate ______________________________

# Chemical equations

# 2 Dissociation and precipitation (cont.)

## Precipitation of ionic compounds




### Recognising precipitation

- When two solutions of different ionic compounds are mixed together, an insoluble ionic compound may form. The insoluble compound 'precipitates' (sinks to the bottom of the container).

### Predicting precipitation

1. Write the ions, in the correct ratio, produced by dissociation in each solution. For example, in the illustration above:
   - solution A contains $K^+ : I^-$
   - solution B contains $Pb^{2+}: 2\ (NO_3)^-$
2. List the new combinations of ions which are possible when the solutions are mixed. For example:
   - $K^+$ and $(NO_3)^-$
   - $Pb^{2+}$ and $I^-$
3. Refer to the solubility rules (see box) to decide if any of the new combinations is contained in an insoluble substance. For example:
   - $K^+$ and $(NO_3)^-$ are contained in $K(NO_3)$ which is soluble.
   - $Pb^{2+}$ and $I^-$ are contained in $PbI_2$ which is insoluble.

*Prediction:* Lead iodide will precipitate. Dissociated potassium and nitrate ions will remain in solution.

### Solubility rules for common salts

- The rules must be applied in order: Rule 1, then Rule 2, etc.
- Substances with solubilities $\leq 10^{-4}$ g/100 g are said to be *insoluble.*

1. Sodium ($Na^+$), potassium ($K^+$) and ammonium ($NH_4^+$) salts are *soluble.*
2. Nitrates and acetates are *soluble.*
3. Silver ($Ag^+$), lead ($Pb^{2+}$), mercury ($Hg^+$) and copper ($Cu^+$) compounds are *insoluble.*
4. Chlorides, bromides and iodides are *soluble.*
5. Carbonates, sulfides, oxides and hydroxides are *insoluble.*
6. Sulfates are *soluble except* calcium sulfate and barium sulfate.

### General precipitation equations

- These equations show the reactants and products as chemical compounds. Subscripts indicate dissolved (aq) or solid (s) compounds. The precipitation equation for the reaction illustrated above is:

$$Pb(NO_3)_{(aq)} + 2\ KI_{(aq)} \rightarrow 2\ K(NO_3)_{(aq)} + PbI_{2(s)}\downarrow$$

The symbol ↓ may be included to emphasise that the precipitate is deposited in the container

## Net ionic precipitation equations

- These equations show only the ions involved in the formation of the precipitate. To write net ionic precipitation equations:
  1. begin with the precipitation equation for the reaction as shown in the example above.
  2. rewrite the precipitation equation to show the reactants and the dissolved product as dissociated ions.
  3. cancel the ions which are identical and have equal numbers on both sides of the equation. These are the spectator ions; they have not been changed by the precipitation reaction. For example:

$$Pb^{2+}_{(aq)} + \cancel{2\ (NO_3)^-_{(aq)}} + \cancel{2\ K^+_{(aq)}} + 2\ I^-_{(aq)} \rightarrow \cancel{2\ K^+_{(aq)}} + \cancel{2\ (NO_3)^-_{(aq)}} + PbI_{2(s)}$$

The net ionic precipitation equation is: $Pb^{2+}_{(aq)} + 2\ I^-_{(aq)} \rightarrow PbI_{2(s)}\downarrow$

**3** For each of the following pairs of solutions, predict whether or not a precipitate will form when the two solutions are mixed. Write balanced net ionic equations for the precipitation reactions occurring.

(a) silver nitrate solution + potassium bromide solution
(b) ammonium sulfate solution + calcium chloride solution
(c) calcium nitrate solution + magnesium chloride solution
(d) sodium sulfide solution + lead nitrate solution
(e) iron III chloride + sodium sulfate solution

Chemical equations

# 4 Acids, bases and neutralisation

## Acids

- In aqueous solution this group of chemical compounds has the general properties of:
  1. a sour taste
  2. the ability to change the colour of litmus (a vegetable dye) from blue to red
  3. reacting with active metals (e.g. Mg, Zn) to form hydrogen gas
  4. reacting with carbonates to form carbon dioxide gas

### Examples of common acids

- *Inorganic acids* are made from chemicals found in the atmosphere or in rocks:
  sulfuric acid $H_2(SO_4)_{(aq)}$
  nitric acid $H(NO_3)_{(aq)}$
  hydrochloric acid $HCl_{(aq)}$
  hydrofluoric acid $HF_{(aq)}$
  phosphoric acid $H_3(PO_4)_{(aq)}$
  carbonic acid $H_2(CO_3)_{(aq)}$
- *Organic acids* (carboxylic acids) are found in living things. They contain the carboxyl group — COOH:
  ethanoic acid (acetic acid) $CH_3COOH_{(aq)}$
  methanoic acid (formic acid) $HCOOH_{(aq)}$
  lactic acid $CH_3CH(OH)COOH_{(aq)}$
  citric acid $HOOCCH_2C(OH)(COOH)CH_2COOH_{(aq)}$
- In the formulas for inorganic acids given above, the radicals have been shown in brackets. This is not necessary in these particular formulas. Students learning to write chemical equations will find it helpful to retain the brackets at all times. Brackets are necessary when the ratio indicates more than one radical (e.g. $Al_2(SO_4)_3$).

### Acid + metal reactions

- Acids contain hydrogen. Active metals react with most acids to displace the hydrogen and produce a salt. For example:

$$Mg_{(s)} + H_2SO_{4(aq)} \rightarrow MgSO_{4(aq)} + H_{2(g)}$$ The salt is magnesium sulfate

$$Mg_{(s)} + 2\ HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$$ The salt is magnesium chloride

- When metals react with organic acids, the hydrogen in the carboxyl group is displaced. For example:

$$Mg_{(s)} + 2\ HCOOH_{(aq)} \rightarrow (HCOO)_2Mg_{(aq)} + H_{2(g)}$$ The salt is magnesium methanoate

**1** Write balanced equations for the reaction of zinc metal with each of the following acids:
(a) dilute sulfuric acid
(b) hydrochloric acid
(c) phosphoric acid
(d) ethanoic acid

**2** Nitric acid and concentrated sulfuric acid do not produce hydrogen with metals. Balance the following equations:
(a) $Cu_{(s)} + HNO_{3(aq)}(\text{concentrated}) \rightarrow Cu(NO_3)_{2(aq)} + NO_{2(g)} + H_2O_{(l)}$
(b) $Cu_{(s)} + HNO_{3(aq)}(\text{dilute}) \rightarrow Cu(NO_3)_{2(aq)} + NO_{(g)} + H_2O_{(l)}$
(c) $Zn_{(s)} + H_2SO_{4(aq)}(\text{concentrated}) \rightarrow ZnSO_{4(aq)} + SO_{2(g)} + H_2O_{(l)}$

### Acid + carbonate reactions

- Acids react with carbonates, either solid or dissolved in water, to form a salt, carbon dioxide and water. For example:

$$Na_2CO_{3(s)} + 2\ HCl_{(aq)} \rightarrow 2\ NaCl_{(aq)} + CO_{2(g)} + H_2O_{(l)}$$ The salt is sodium chloride

**3** (a) Write balanced equations for the reaction of solid sodium carbonate with each of the following acids:
(i) sulfuric acid
(ii) nitric acid
(iii) phosphoric acid

(b) Marble, a rock, is composed mainly of calcium carbonate. A piece of rock can be identified as marble if fizzing occurs when a small amount of hydrochloric acid is added to the rock.
(i) Name the gas produced in the fizzing. ______________________
(ii) Write a balanced equation for the reaction occurring.
______________________

2

The mole

# Mass ratio and mole ratio [cont.]

## Molar mass from *n* and mass of sample

- To find the mass of 1 mole of a substance, given the number of moles and the mass of the sample, rearrange the mathematical formula:

$$\text{Number of moles, } n = \frac{\text{Mass}}{\text{MM}}$$

$$\therefore \text{Mass} = \text{MM} \times n \qquad \textit{Note}\text{: Mass must be in grams.}$$

*Example:* 0.1 mole of a gas has a mass of 2.6036 g. Find the molar mass of the gas.

$$\text{MM} = \frac{\text{Mass}}{n} = \frac{2.6036\ \text{g}}{0.1} = 26.036\ \text{g}$$

**3** Find the molar mass of the substance in each of the following:

(a) 3 moles have a mass of 210.39 g.
Molar mass = ________________

(b) The mass of 0.2 moles is 5.6104 g.
Molar mass = ________________

(c) The mass of 0.05 mole is 5.2995 g.
Molar mass = ________________

(d) 77.9675 g contain 0.25 mole.
Molar mass = ________________

(e) A 10 mole sample has mass 0.34 kg.
Molar mass = ________________

(f) 0.12 mole has a mass of 8892 mg.
Molar mass = ________________

(g) 600 mg contain $1.5 \times 10^{-2}$ mole.
Molar mass = ________________

(h) $3 \times 10^{-2}$ mole has a mass of 4.74 g.
Molar mass = ________________

(i) 1.25 mole has a mass of 60 g.
Molar mass = ________________

(j) The mass of 1/2 mole is 116.7 g.
Molar mass = ________________

## Problems involving equations

Step 1: Write a balanced equation.
Step 2: Substitute molar mass values in the equation.
Step 3: Calculate for given masses.

*Example:* What mass of MgO would form if 0.24 g of Mg were completely burned?

Step 1: $2\ Mg_{(s)} + O_{2(g)} \rightarrow 2\ MgO_{(s)}$
Step 2: $2 \times 24.31\ g + 2 \times 16.00\ g \rightarrow 2\ (24.31 + 16.00)$
$48.62\ g + 32.00\ g \rightarrow 80.62\ g$
Step 3: $0.24\ g \rightarrow ?\ g\ MgO$

$$\text{grams MgO} = \frac{0.32 \times 80.62}{48.62} = 0.53\ \text{g}$$

**4**

(a) Calculate the mass of calcium oxide produced if 0.2 g calcium were completely reacted with oxygen gas.
(b) A 0.3 gram sample of sodium metal was burned completely in chlorine gas. Calculate the mass of sodium chloride formed.
(c) What mass of copper II hydroxide would be produced by the complete reaction of 4.236 g of copper with sodium hydroxide solution?
(d) How many grams of magnesium metal would be needed to produce 12.274 g of magnesium bromide?
(e) What mass of aluminium is needed to make 3.0588 g of aluminium oxide?