Assumed Knowledge & Skills

The Chemistry Stage 2 Curriculum Statement assumes that students will have completed at least one year’s study of chemistry. The following lists should not be regarded as either a course or a teaching sequence. Although students may not need to be acquainted with all the following skills and concepts, they should be familiar with most.

The numbers in parentheses indicate specific curriculum statement topics to which the skills and knowledge of chemistry relate.

Skills

- Common laboratory equipment, including heating apparatus, electronic balance, measuring cylinder, and thermometer, should be used with care and safety. (1)
- Laboratory instructions should be understood and followed. (1)
- Filtration, evaporation, and decanting are common laboratory procedures. (1)
- A hypothesis is a proposal that can be tested experimentally. (1)
- A controlled experiment is designed to test a hypothesis. It requires one variable (the independent variable) to be manipulated while another (the dependent variable) is measured. Other variables must be controlled. (1)
- Experimental results must be observed and recorded accurately. (1)
- Tables are often a convenient form for the recording of data. Such a table should have column headings and units showing clearly what is recorded. (1)
- Data can frequently be better presented graphically with the independent variable plotted on the horizontal axis and the dependent variable plotted on the vertical axis. Axes should have regular scales and be clearly labelled with the units shown. The relationship between the variables is shown by a line of best fit. (1)
- Observations are the changes seen or measured in an experiment. Conclusions are inferred from such observations. (1)
- There are sources of error in all experiments. (1)
- Locating information involves finding, selecting, and noting relevant materials. (1, 2, 3, 4, 5, 6)
- Effective communication of scientific information requires concise, unambiguous expression with correct grammar and spelling and the correct use of scientific terms and conventions. (1, 2, 3, 4, 5, 6)

Knowledge of Chemistry

Matter

- Matter exists as particles. (2)
- Electric current is a movement of charged particles: valence electrons in metals and ions in ionic melts or solutions. (2, 4)
- In solids the particles are closely spaced in fixed positions. In liquids the particles are closely spaced but able to move about. In gases the particles are far apart and free to move about. (2)
- Matter can be divided into elements, compounds, and mixtures. (2)
- Elements can be conveniently divided into metals that are good electrical conductors and non-metals that are poor electrical conductors. (2)
- The properties of an element or a compound are related to its structure (the arrangement of the particles) and bonding (the forces between the particles). (2)
- On the basis of its structure, matter can be divided into molecular and non-molecular. (2)
- In non-molecular substances there are only strong forces of attraction between the particles; these are known as primary bonds. (2)
- The bonding in non-molecular substances can be metallic, covalent, or ionic. (2)
- Attractive forces exist between bodies with opposite charges; repulsive forces exist between bodies with like charges. (2, 6)
All matter is composed of atoms, which consist of a nucleus that contains protons and neutrons and is surrounded by a cloud of electrons. (2)

The electrons in an atom are arranged in energy levels. (2, 3)

In substances with metallic bonding the valence electrons are able to move freely from one atom to another. (2)

Non-metallic atoms bond to each other by sharing valence shell electrons in covalent bonds; this often gives each atom eight electrons in its valence shell. (2)

When metallic atoms react with non-metallic atoms the metallic atoms lose valence electrons and form positively charged ions, whereas the non-metallic atoms gain electrons and form negatively charged ions. (2)

There is a strong force of attraction between the oppositely charged ions. (2)

Group 1 and Group 2 metals form 1+ and 2+ ions respectively, whereas oxygen and sulfur form 2− ions and the halogens 1− ions. (2)

In molecular substances there are weak forces of attraction between the molecules; these are known as interactions. (2, 3, 5, 6)

The strength of interactions depends on the size and polarity of the molecules involved. (2, 3, 5, 6)

Repulsion between bonding and non-bonding pairs of valence shell electrons determines the spatial arrangement of bonded atoms. (2)

The periodic table is a way of displaying the elements so that similarities and trends in properties are more apparent. (2)

The position of an element in the periodic table is related to its electron configuration. (2)

The position of an element in the periodic table is related to its metallic or non-metallic character. (2)

Electronegativity is the tendency of an element to gain electrons; metals have low electronegativities whereas non-metals have higher electronegativities. (2)

The electronegativity of elements increases across and decreases down the periodic table. (2)

The charges of the monatomic ions formed by the elements of atomic numbers 1 to 20 are related to the number of electrons in the outside (valence) shell. (2)

The number of covalent bonds formed by the elements of atomic numbers 1 to 20 is related to the number of electrons in the outside (valence) shell. (2)

Non-metallic atoms are able to bond covalently to form ions such as OH−, NO3−, SO42−, CO32−, and NH4+. (2)

The metals iron, zinc, copper, silver, and lead commonly occur in their compounds as Fe2+, Fe3+, Zn2+, Cu2+, Ag+, and Pb2+ ions. (2)

In covalent bonds between atoms of different electronegativity the electrons are shared unequally; such a bond is known as a ‘polar bond’. (2, 3, 5)

Heat is a form of energy. (4)

Energy can be transferred by radiation. Radiant energy can be absorbed, reflected, or transmitted by matter. (3)

Temperature is a measure of the average kinetic energy of the particles present, whereas heat is a measure of the total amount of kinetic energy present. (4)

Air is a mixture of gases: nitrogen (78% by volume), oxygen (21%), argon (0.9%), and carbon dioxide (0.03%). (2)

Reactions

A chemical change involves the formation of a new substance and is accompanied by the gain or loss of energy. (4)

Chemical equations can be written to describe a chemical change. (2, 3, 4, 5, 6)

Ionic equations that omit non-reacting species, commonly known as ‘spectator ions’, can be written. (2, 3, 4, 5, 6)

When a substance burns in air it combines with oxygen. (2)

Salts containing sodium, potassium, ammonium, or nitrate ions are soluble in water. (2, 3)

If an ionic substance dissolves in water, the ions dissociate with each ion surrounded by water molecules. (2, 6)

An insoluble ionic substance forms as a precipitate when solutions containing the ions are mixed. (4, 6)

Self-ionisation occurs in water to a small degree, giving rise to low concentrations of OH− and H3O+ (H+) ions. (3)

Acids are compounds or ions that donate protons, whereas bases are compounds or ions that accept protons. (2, 5)

Common acids include hydrochloric acid (HCl), nitric acid (HNO3), and sulfuric acid (H2SO4). (2, 3, 4, 5)
Acids react with metal oxides and hydroxides to produce a salt and water, and with carbonates to produce a salt and water, and carbon dioxide. (2, 3, 4)

An aqueous solution can be described as acidic, alkaline (basic), or neutral on the basis of the concentration of H⁺ ions in the solution. (2, 5, 6)

An aqueous solution can be described as acidic, alkaline (basic), or neutral on the basis of its pH. (2, 5, 6)

Ammonia is a base. (2, 5, 6)

Oxidation is defined as the gain of oxygen, the loss of electrons, or the increase in oxidation number. (3, 4, 5)

Reduction is defined as the loss of oxygen, the gain of electrons, or the decrease in oxidation number. (3, 4, 5)

Oxidation and reduction are complementary processes. (3, 4, 5)

When electrons are transferred in a reaction, half-equations can be written for the gain and the loss of electrons; full equations can be written by combining two half-equations. (3, 4, 5)

Metals differ in their tendency to lose electrons; more reactive metals lose electrons more easily. (4)

Potassium, sodium, and calcium react with water to form hydrogen gas, hydroxide ions, and ions of the metal. (4)

Magnesium, zinc, and iron react with dilute acids to form hydrogen and salts of the metals. (4)

If a metal readily loses electrons (is readily oxidised), its ion does not readily gain electrons (is not easily reduced). (4)

A more reactive metal is able to donate electrons to the ion of a less active metal. (4)

Oxidation number is a useful means of identifying redox reactions. (4)

Carbon Chemistry

Organic chemistry is the chemistry of carbon compounds. (5)

Hydrocarbons are compounds that contain only carbon and hydrogen. (4, 5)

Hydrocarbons are commonly used as fuels, the products of complete combustion being water and carbon dioxide. (4, 5)

The molecular formula shows the atoms present in a molecule of a substance. The structural formula of a compound shows unambiguously the arrangement of the atoms. (2, 5)

Alkanes and alkenes are named systematically. (5)

A functional group is the reactive part of an organic molecule. (5)

Compounds with the same functional group undergo similar reactions. (5)

The alkyl component of an organic compound is generally unchanged in an organic reaction. (5)

Isomers are different compounds that have the same molecular formula. (5)

A benzene ring occurs in many compounds; it is a reasonably unreactive group. (5)

Compounds existing as small molecules are likely to be more soluble than larger molecules of a similar nature. (5)

Chemical Calculations

The quantities of different substances can be conveniently compared by the use of the mole as a unit. (3)

The molar mass of a substance can be derived by the addition of the relative atomic masses of the elements present, with the answer expressed in grams. (3)

The amount of a substance (in moles) is related to the mass, \( m \) (in grams), and the molar mass, \( M \) (in g mol⁻¹). (3)

The molar concentration (or molarity), \( c \) (in mol L⁻¹), of a solution is related to the amount of solute, \( n \) (in moles), and the volume of the solution, \( V \) (in litres). (3)

The concentration of a solution can be related to the mass of solute (in grams) and the volume of the solution (in litres). (3)

The relative amounts (in moles) of substances reacted or produced during a reaction are indicated by the coefficients in the balanced equation for the reaction. (3)

Given the equation for a reaction, the quantity of one reactant or product involved in a chemical reaction can be used to determine the quantity of another. (3)

Numerical answers are limited by the least accurate data used in the calculation. (3)

Large and small quantities are more conveniently expressed by means of scientific notation. (3)

Calculators frequently display scientific notation in different ways. (3)

Substances vary in density, which is the mass of material per unit volume. (3)
Molecular Structures

There are a number of methods for showing the arrangements of atoms in molecules. Students could expect to encounter a number of these in textbooks and other reading as well as in tests and examinations.

An understanding of the different shapes of molecules needs to be developed over a period of time and requires an understanding of covalence of the atoms involved. A first step towards developing an ability to draw and interpret molecular structures could be the use of electron dot formulae (also known as electronic formulae in some books) to allow students to appreciate the number of electrons in the valence shells of the atoms involved.

The next step from such a diagram could involve the use of lines to represent pairs of electrons. These are sometimes referred to as bond diagrams. Such diagrams should show the spatial arrangement of the atoms:

In organic chemistry the function of a diagram, known as a structural formula, is to show unambiguously the arrangement of the atoms to allow distinction to be made between isomers. Structural formulae do not attempt to show the spatial arrangement of the atoms.

A structural formula showing each atom separately and each electron pair as a line is known as an extended or expanded structural formula. Such formulae are very useful for smaller molecules but become extremely confusing for formulae containing more than three carbon atoms.

When students have mastered these structural formulae they would be ready to move on to more abbreviated forms, two of which are shown below. While students would be expected to be able to interpret both of the following types of structural formulae they should be able to choose any format for drawing them.

1. **Condensed structural formula**, showing groups in the molecule, with lines showing the covalent bonds between the groups. In drawing a bond between the main chain and a side group or functional group, the line drawn should join the atoms involved in the bond.

2. **Skeleton structural formula**, omitting hydrogen atoms except those in functional groups. Such formulae are the most convenient for large molecules, such as carbohydrates.

Another means of drawing structural formulae is common in many textbooks:

This type of structural formula can be more difficult to interpret and, if used by students, commonly is drawn incorrectly. This type of structural formula will not be used in examinations although students may use it if they wish.
Balancing Equations

Chemical Equations

- In a chemical reaction, the way in which the atoms are joined together is changed. Bonds are broken and new ones are formed as the **reactants** are changed into **products**. The same atoms (both number and type) are present before and after the reaction — they are just arranged differently.

![Chemical Reaction Diagram]

- The physical state of the substances should be added to the equation. These are shown in brackets after each symbol or formula: (s) for a solid, (l) for a liquid, (g) for a gas, and (aq) for a solution of water.

<table>
<thead>
<tr>
<th>Symbol</th>
<th>Meaning</th>
</tr>
</thead>
<tbody>
<tr>
<td>(s)</td>
<td>a reactant or product in the solid state</td>
</tr>
<tr>
<td>(l)</td>
<td>a reactant or product in the liquid state</td>
</tr>
<tr>
<td>(aq)</td>
<td>an aqueous solution (the substance is dissolved in water)</td>
</tr>
<tr>
<td>(g)</td>
<td>a reactant or product in the gaseous state</td>
</tr>
<tr>
<td>Heat →</td>
<td>indicates that heat is supplied to the reaction</td>
</tr>
<tr>
<td>Pt →</td>
<td>a formula written above the arrow indicates its use as a catalyst (in this case, platinum)</td>
</tr>
</tbody>
</table>

- When a reaction requires heat, the word ‘heat’ is written above the arrow to denote that heat has been applied.
Balancing Chemical Equations

- To represent chemical reactions correctly, equations must be *balanced*. This means the number of atoms of each element must be the same on both sides of the equation.

- According to the **law of conservation of mass**, the total mass of the reactants in a chemical reaction is equal to the total mass of the products. Atoms are not created or destroyed, but are rearranged to form new substances. In order to balance an equation, numbers called **coefficients** are placed *in front* of the whole formulas.

Rules for Balancing Equations

1. Write the reactants and products using the correct formula and state for each substance.

2. Count the number of atoms of each element on the left-hand side of the equation. Do the same for the right-hand side and compare for each element. If any of these numbers do not match, the equation is not balanced and you will need to proceed to the following steps.

3. Balance by placing coefficients in front of the formulae. Do not change the actual formula. If any substance is present as an element, leave the balancing of it to last.

4. Check all atoms or ions to ensure that they are balanced.

5. Make sure that the coefficients are in their lowest possible ratio.

- When balancing equations, remember:
  
  - You cannot change the small subscript numbers in a formula. You can only add whole numbers in front of a formula.
  
  - Atoms cannot appear from nowhere, nor can they disappear into thin air; this is why equations must be balanced.
  
  - If you place a number in front of a compound like \( \text{Fe}_2(\text{SO}_4)_3 \), you have to multiply all the atoms in the formula by that number.
  
  - The subscripts (s), (g), (l) and (aq) should be used to show the physical state of reactants and products.
Examples:

**Solution:**

**Step 1** Write the reactants and products including states.  
\[ \text{NH}_3(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{NO}(\text{g}) + \text{H}_2\text{O}(\text{g}) \]

**Step 2** This is an unbalanced equation because there are three atoms of hydrogen on the left-hand side of the equation and only two atoms of hydrogen on the right-hand side.

**Step 3** Begin to balance the equation by correcting for hydrogen atoms. A coefficient of 2 for \( \text{NH}_3 \) and a coefficient of 3 for \( \text{H}_2\text{O} \) will give six atoms of hydrogen on both sides:  
\[ 2\text{NH}_3(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{NO}(\text{g}) + 3\text{H}_2\text{O}(\text{g}) \]

Now, an imbalance in nitrogen atoms has occurred. Balance the nitrogen atoms by introducing a coefficient of 2 for \( \text{NO} \).  
\[ 2\text{NH}_3(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{NO}(\text{g}) + 3\text{H}_2\text{O}(\text{g}) \]

**Step 4** Now there are two atoms of oxygen on the left and five on the right. Balance the oxygen atoms. This can be done by introducing a coefficient of \( \frac{1}{2} \) for \( \text{O}_2 \).  
\[ 2\text{NH}_3(\text{g}) + 2\frac{1}{2} \text{O}_2(\text{g}) \rightarrow 2\text{NO}(\text{g}) + 3\text{H}_2\text{O}(\text{g}) \]

**Step 5** Check the equation. The same numbers of N, H and O atoms are on both sides of the equation, and states have been included, but the ratios are not yet in the simplest ratio. The equation is balanced.

**Step 6** Since the usual practice is to have whole-number coefficients, multiply the entire equation by 2.  
\[ 4\text{NH}_3(\text{g}) + 5\text{O}_2(\text{g}) \rightarrow 4\text{NO}(\text{g}) + 6\text{H}_2\text{O}(\text{g}) \]
Further Questions

1. Balance the following equations.

   (a) $\text{N}_2(g) + \text{H}_2(g) \rightarrow \text{NH}_3(g)$

   (b) $\text{SiO}_2(s) + \text{C}(s) \rightarrow \text{Si}(s) + \text{CO}_2(g)$

   (c) $\text{FeO}(s) + \text{O}_2(g) \rightarrow \text{Fe}_2\text{O}_3(s)$

   (d) $\text{Cr}(s) + \text{S}_8(s) \rightarrow \text{Cr}_2\text{S}_3(s)$

   (e) $\text{NaHCO}_3(s) \rightarrow \text{Na}_2\text{CO}_3(s) + \text{CO}_2(g) + \text{H}_2\text{O}(g)$

2. Balance the following equations.

   (a) $\text{HCl}_{(aq)} + \text{Al}_{(s)} \rightarrow \text{AlCl}_3(aq) + \text{H}_2(g)$

   (b) $\text{NH}_3(g) + \text{O}_2(g) \rightarrow \text{NO}(g) + \text{H}_2\text{O}(g)$

   (c) $\text{NO}(g) + \text{O}_2(g) \rightarrow \text{NO}_2(g)$

   (d) $\text{Na}_{(s)} + \text{P}_4(s) \rightarrow \text{Na}_3\text{P}_{(s)}$
3. Balance the following equations.

(a) \( \text{Cu} + \text{O}_2 \rightarrow \text{CuO} \)

(b) \( \text{HgO} \rightarrow \text{Hg} + \text{O}_2 \)

(c) \( \text{AsCl}_3 + \text{H}_2\text{S} \rightarrow \text{As}_2\text{S}_3 + \text{HCl} \)

(d) \( \text{Fe}_2\text{O}_3 + \text{H}_2 \rightarrow \text{Fe} + \text{H}_2\text{O} \)

(e) \( \text{NaCl} \rightarrow \text{Na} + \text{Cl}_2 \)

(f) \( \text{Al} + \text{H}_2\text{SO}_4 \rightarrow \text{H}_2 + \text{Al}_2(\text{SO}_4)_3 \)

(g) \( \text{C}_8\text{H}_{18} + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O} \)
<table>
<thead>
<tr>
<th>Cations</th>
<th>+1</th>
<th>+2</th>
<th>+3</th>
</tr>
</thead>
<tbody>
<tr>
<td>lithium</td>
<td>Li⁺</td>
<td>Mg²⁺</td>
<td>Al³⁺</td>
</tr>
<tr>
<td>sodium</td>
<td>Na⁺</td>
<td>Ca²⁺</td>
<td>Cr³⁺</td>
</tr>
<tr>
<td>potassium</td>
<td>K⁺</td>
<td>Ba²⁺</td>
<td>Fe³⁺</td>
</tr>
<tr>
<td>caesium</td>
<td>Cs⁺</td>
<td>Fe⁴⁺</td>
<td>Cl⁻</td>
</tr>
<tr>
<td>silver</td>
<td>Ag⁺</td>
<td>Ni²⁺</td>
<td>Hg²⁺</td>
</tr>
<tr>
<td>copper(I)</td>
<td>Cu⁺</td>
<td>Zn²⁺</td>
<td>Mn²⁺</td>
</tr>
<tr>
<td>ammonium</td>
<td>NH₄⁺</td>
<td>Pb²⁺</td>
<td>Sr²⁺</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Anions</th>
<th>−1</th>
<th>−2</th>
<th>−3</th>
</tr>
</thead>
<tbody>
<tr>
<td>hydride</td>
<td>H⁻</td>
<td>O²⁻</td>
<td>N³⁻</td>
</tr>
<tr>
<td>fluoride</td>
<td>F⁻</td>
<td>S²⁻</td>
<td>P³⁻</td>
</tr>
<tr>
<td>chloride</td>
<td>Cl⁻</td>
<td>SO₄²⁻</td>
<td>PO₄³⁻</td>
</tr>
<tr>
<td>bromide</td>
<td>Br⁻</td>
<td>CO₃²⁻</td>
<td></td>
</tr>
<tr>
<td>iodide</td>
<td>I⁻</td>
<td>S²⁻</td>
<td></td>
</tr>
<tr>
<td>hydroxide</td>
<td>OH⁻</td>
<td>HCO₃⁻</td>
<td></td>
</tr>
<tr>
<td>nitrate</td>
<td>NO₃⁻</td>
<td>SO₄²⁻</td>
<td></td>
</tr>
<tr>
<td>hydrogen carbonate</td>
<td>HCO₃⁻</td>
<td>hydrogen phosphate</td>
<td></td>
</tr>
<tr>
<td>hydrogen sulfate</td>
<td>HSO₄⁻</td>
<td>HPO₄²⁻</td>
<td></td>
</tr>
<tr>
<td>chlorate</td>
<td>ClO₄⁻</td>
<td></td>
<td></td>
</tr>
<tr>
<td>hydrogen sulfite</td>
<td>HSO₃⁻</td>
<td></td>
<td></td>
</tr>
<tr>
<td>nitrite</td>
<td>NO₂⁻</td>
<td></td>
<td></td>
</tr>
<tr>
<td>permanganate</td>
<td>MnO₄⁻</td>
<td></td>
<td></td>
</tr>
<tr>
<td>hypochlorite</td>
<td>OCl⁻</td>
<td></td>
<td></td>
</tr>
<tr>
<td>dihydrogen phosphate</td>
<td>H₂PO₄⁻</td>
<td></td>
<td></td>
</tr>
<tr>
<td>cyanide</td>
<td>CN⁻</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
Ionic Compounds & Solubility

Solubility

- **Solubility**: the solubility of one substance in another is defined as the extent to which a solute can dissolve in a solvent. Dissolving occurs when one substance is pulled a part by another, or moves into the spaces between another.

  \[
  \text{solute} + \text{solvent} = \text{solution}
  \]

- Solubility is determined by the polarity of the solute and solvent.

  ‘like dissolves like’

- **Unsaturated solution**: less than the maximum amount of solute has been added to the solvent.
- **Saturated solution**: contains the maximum amount of solute for the volume of solution at a given temperature.
- **Super saturated solution**: contains more solute than it should at that temperature. The solute may begin to crystallise as the temperature is decreased.

- Water is an excellent solvent. Water samples containing dissolved substances are called aqueous solutions.

Dissociation of Ionic Compounds

- Ionic compounds contain **anions** and **cations** held in a lattice by electrostatic forces.
- When placed in water, the positive regions of water molecules are attracted to negative ions. The negative regions are attracted to positive ions.
- When dissolved the ions are surrounded by water molecules – the ions are hydrated.
- This process of separation of ions from a lattice is known as **dissociation**.
Precipitation of Ionic Compounds

- When two solutions of different ionic compounds are mixed together, an **insoluble** ionic compound may form. The insoluble compound **precipitates.**

<table>
<thead>
<tr>
<th>Name of ion</th>
<th>Symbol</th>
<th>Soluble compounds of ion</th>
<th>Insoluble compounds of ion</th>
</tr>
</thead>
<tbody>
<tr>
<td>group I ions</td>
<td>Li⁺, Na⁺, K⁺, Rb⁺, Cs⁺, Fr⁺</td>
<td>all</td>
<td>none</td>
</tr>
<tr>
<td>ammonium</td>
<td>NH₄⁺</td>
<td></td>
<td></td>
</tr>
<tr>
<td>hydrogen</td>
<td>H⁺</td>
<td></td>
<td></td>
</tr>
<tr>
<td>nitrate</td>
<td>NO₃⁻</td>
<td></td>
<td></td>
</tr>
<tr>
<td>nitrite</td>
<td>NO₂⁻</td>
<td></td>
<td></td>
</tr>
<tr>
<td>chlorides</td>
<td>Cl⁻</td>
<td></td>
<td>Ag⁺, Pb²⁺, Hg²⁺ (PbCl₂ is moderately soluble in hot water.)</td>
</tr>
<tr>
<td>bromides</td>
<td>Br⁻</td>
<td></td>
<td>Ba²⁺, Pb²⁺ (Ag₂SO₄ and CaSO₄ are slightly soluble.)</td>
</tr>
<tr>
<td>iodides</td>
<td>I⁻</td>
<td></td>
<td></td>
</tr>
<tr>
<td>sulfates</td>
<td>SO₄²⁻</td>
<td></td>
<td></td>
</tr>
<tr>
<td>carbonates</td>
<td>CO₃²⁻</td>
<td>Na⁺, K⁺, NH₄⁺</td>
<td></td>
</tr>
<tr>
<td>phosphates</td>
<td>PO₄³⁻</td>
<td></td>
<td></td>
</tr>
<tr>
<td>sulfides</td>
<td>S²⁻</td>
<td>Na⁺, K⁺</td>
<td></td>
</tr>
<tr>
<td>hydroxides</td>
<td>OH⁻</td>
<td>Na⁺, K⁺, Ba³⁺ (NH₄OH and (NH₄)₂O do not exist as solids.)</td>
<td></td>
</tr>
<tr>
<td>oxides</td>
<td>O²⁻</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**Table 11.2** Solubility in water of compounds of common ions
1) Use the particle theory of matter to explain the difference between a solid, liquid and gas.

_________________________________________________________________________
_________________________________________________________________________
_________________________________________________________________________
_________________________________________________________________________
_________________________________________________________________________
_________________________________________________________________________

2) Distinguish between a chemical change and a physical change.

_________________________________________________________________________

3) (a) List five characteristic properties of non-metals.
1
2
3
4
5
(b) List five characteristics of metals.
1
2
3
4
5

4) Explain each of the following and provide a common example for each:
(a) an element?
_________________________________________________________________________
_________________________________________________________________________

(b) a compound?
_________________________________________________________________________
5) Consider the following elements 12A, 28B, 18C and 34D.
Use a periodic table and design a table that give details of:
(a) Atomic number
(b) Electronic configuration
(c) Group and Period
(d) Correct name and symbol
(e) The nature of each element (ie metal, semi-metal or non-metal)

(f) Element B exists in nature in two forms 58B, 60B. Outline the similarities and differences between these two. What name is given to these two forms of element B?

6) Calculate the mass of solute present in each of the following solutions:
(a) 25.0 mL of 5.00 M sodium chloride solution.
(b) 2.00 L of 0.500 M iron (III) chloride solution
7) 50mL of 4.00 M sulfuric acid is diluted to 250 mL. If 25 mL of the diluted solution is taken, how many moles of sulfuric acid are present in the aliquot?

8) What volume of 0.100M KOH solution will neutralise 25.0 mL of 0.300 M sulfuric acid solution?

9) If 50.54 percent of naturally occurring bromine atoms have an Ar of 78.92 and 49.46 percent have an Ar of 80.92, calculate the relative atomic mass of bromine.

10) Using subshell notation, write the electron configuration for the following atoms and ions.
   a) Argon atom: ________________________________
   b) Sodium atom: ________________________________
   c) Sodium ion: ________________________________
   d) Chloride ion: ________________________________