National 5 Chemistry

Unit 1:
Chemical Changes & Structure

Student:

Topic 4: Chemistry of Acids & Bases

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4.1 Acids & Bases

Common Acids

Acids are all around you. Most are safe to handle but some are definitely not.

Fizzy drinks (CO₂, carbonic acid), some fruits like lemon and orange (citric acid), vinegar and ketchup (ethanoic acid), and even many vegetables (ascorbic acid, vitamin C) contain acid. These are examples of weak acids. A more unpleasant example is a bee sting, nettle sting or ant bite (formic acid).

Stronger acids are found in car battery (sulfuric acid) while our stomach contain acid (hydrochloric) to break down our food into smaller molecules.

- **Acids taste sour**
  - the Latin word for so was “acidus”

- **Acids kill cells**
  - the most vulnerable part of you is your eyes, because they have living cells on the surface (Safety Glasses!)

  Pickling food in vinegar (ethanoic acid) is an ancient way of killing bacteria and keeping food. It made food taste sour, but many people like that, so we still pickle food.

- **Acids react with metals**
  - iron objects, like the Forth Rail Bridge, rust much faster these days because of acid rain.

- **Acids react with carbonates**
  - acid rain is destroying many of the marble carbonates (calcium carbonate) statues in the world

- **Acids react with alkalis**
  - the leaves of a dock (or docken) plant can neutralise the sting from a nettle

- **Acids change the colour of indicators**
  - e.g. they turn universal indicator from green to red

- **Acids have a pH less than 7**
  - on a pH scale that goes from below 0 to above 14.
**Common Bases**

**Bases** are also all around you. *Most* are *safe* to handle but some are definitely not.

To **paste** contains a **base** to help neutralise the **acid** on your **tee** , produced by **bac** . Most **soa** and **dete** contain bases to help cope with **gre** and **oi** stains. Our **liv** produces **bile** (a **base**) to help break down **fat** foods. **Far** and **gar** will spread **li** (calcium hydroxide) on the **so** if it is too **aci**.

Wasp stings are basic, and just as painful as any acid. Other harmful **bases** are found in **blea** (ammonia)

- **Bases feel slippy**
  - most **soa** are basic

- **Bases kill cells**
  - the most vulnerable part of you is your **ey**, because they have **liv** cells on the **sur** (Safety Glasses !)

- **Bases react with acids**
  - **wa** stings (**alk**) should be treated with an acid like **vin** or **lem** juice, while **bee** stings need **bak** **soda** (a base).

**Aci** fumes (SO\(_2\) and CO\(_2\)) from **co** burning **pow** stations are passed through **li** to be neutralised.

The **hum** stomach produces **hy chl** acid to help the **enz** (cat ) **pepsin** to break down **pro** . Sometimes too much **ac** is made and it begins to attack the stomach **wa** causing **pa** . All stomach remedies contain **bases** to **neut** the stomach acid.

- **Bases change the colour of indicators**
  - e.g. they turn **univ** indicator from **gr** to **pur**

- **Bases have a pH greater than 7**
  - on a pH scale that goes from below **7** to above .
Our main source of *bases* are the **Metal Oxides**.

e.g.,

- **lime**  
  - *cal*  
  - *oxide*
- **soda**  
  - *sod*  
  - *oxide*
- **magnesia**  
  - *mag*  
  - *oxide*
- **pearl ash**  
  - *pot*  
  - *oxide*

Though, we also use many **Metal Carbonates**.

e.g.,

- **limestone**  
  - *cal*  
  - *carbonate*
- **marble**  
  - *cal*  
  - *carbonate*
- **baking soda**  
  - *sod*  
  - *carbonate*
- **potash**  
  - *pot*  
  - *carbonate*

All of these *bases* are ionic compounds and, therefore, *so* at room temperature. As *so* , they can be directly added to an acid and will *neu* the acid.

<table>
<thead>
<tr>
<th>base</th>
<th>+</th>
<th>acid</th>
<th>→</th>
<th>salt</th>
<th>+</th>
<th>water</th>
</tr>
</thead>
</table>

Sometimes, however, we prefer to use *sol* of soluble *bases* which we then can call *alk* .

Most *alk* are made by *dis* a me oxide in wa - though only those in **Group 1** are very *sol* . *(Data Booklet).* These oxides react with water to produce hyd ions, \( \text{OH}^- \) (aq).

\[
\begin{align*}
\text{soda} & \quad + \quad \text{H}_2\text{O} & \rightarrow & \text{NaOH} & \quad \text{(aq)} \\
\text{K}_2\text{O} & \quad + \quad \text{H}_2\text{O} & \rightarrow & \text{KOH} & \quad \text{(aq)} \\
\text{CaO} & \quad + \quad \text{H}_2\text{O} & \rightarrow & \text{Ca(OH)}_2 & \quad \text{(aq)} \\
\text{magnesia} & \quad + \quad \text{H}_2\text{O} & \rightarrow & \text{Mg(OH)}_2 & \quad \text{(aq)}
\end{align*}
\]

**The soluble metal oxides can form alkanis with water**
Making Acids

Probably one of the first times the word ac was met was in the phrase “ac ra”.

Gas like nitro dioxide, NO₂, (produced in pet engines) and sul dioxide, SO₂, (co burning pow stations) both dissolve to produce ac solu.

\[
\text{SO}_2 (g) + \text{H}_2\text{O} (l) \rightarrow \text{H}_2\text{SO}_3 (aq)
\]

Fiz drinks are ac because car dioxide, CO₂, dissolves to form carb acid, H₂CO₃.

Other n -met oxi like SO₃ (sulf acid, H₂SO₄) and P₂O₅ (phosph acid) behave this way, though insol oxi like carbon mo ide, CO, cannot form ac solutions.

The soluble non-metal oxides can form acids with water

\[
\begin{align*}
\text{SO}_2 (g) + \text{H}_2\text{O} (l) & \rightarrow \text{H}_2\text{SO}_3 (aq) \\
\text{SO}_3 (g) + \text{H}_2\text{O} (l) & \rightarrow \text{H}_2\text{SO}_4 (aq) \\
\text{CO}_2 (g) + \text{H}_2\text{O} (l) & \rightarrow \text{H}_2\text{CO}_3 (aq) \\
\text{NO}_2 (g) + \text{H}_2\text{O} (l) + \text{O}_2 (g) & \rightarrow \text{H}_2\text{NO}_3 (aq) \\
\text{P}_2\text{O}_5 (s) + \text{H}_2\text{O} (l) & \rightarrow \text{H}_3\text{PO}_4 (aq)
\end{align*}
\]

Not all of our acids are made from oxides, however.

\[
\begin{align*}
\text{H}_2\text{S} (g) + \text{H}_2\text{S} (s) & \rightarrow \text{H}_2\text{S}_2 \text{O}_3 (aq)
\end{align*}
\]

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Q1. Int2
Acids and Bases are found everywhere around us. For each of the substances listed below, decide whether they are acids (A) or bases (B).

lemon juice  toothpaste  lime
wasp stings vitamin C  Coca Cola
tomato ketchup  nettle sting  bleaches
baking soda  stomach juices  detergents

Q2. S
The grid shows the formulae of some oxides.

A \text{ZnO} \quad \text{B} \quad \text{NO}_3 \quad \text{C} \quad \text{K}_2\text{O}
\text{D} \quad \text{CaO} \quad \text{E} \quad \text{Fe}_2\text{O}_3 \quad \text{F} \quad \text{CO}

a) Identify the two oxides which are covalent.

b) Identify the oxide which dissolves in water to give an alkaline solution. You may wish to use the data booklet to help you.

Q3. N5
Which of the following compounds is a base?

A Sodium carbonate
B Sodium chloride
C Sodium nitrate
D Sodium sulfate

Q4. N5
An element was burned in air. The product was added to water, producing a solution with a pH less than 7. The element could be

A tin
B zinc
C sulfur
D sodium

Q5. SC
The grid shows the names of some oxides.

A Magnesium oxide  B Sulphur dioxide  C Copper(II) oxide
D Sodium oxide  E Silicon dioxide  F Calcium oxide
G Iron(III) oxide  H Potassium oxide  I Carbon dioxide

a) Give the four boxes in the grid containing chemicals that would make an alkaline solution.

b) Give the two boxes in the grid containing chemicals that would make an acidic solution.

c) Give the three boxes in the grid containing chemicals that would make neither an acidic nor an alkaline solution.

Q6. Int2
The grid shows pH numbers and test colours of some solutions.

A Purple  B pH 5  C pH 7
D pH 1  E Orange  F pH 14
G pH 8  H Red  I pH 4

a) Which box or boxes show acid pH colours

b) Give the two boxes with alkali pH numbers.

c) Hydrochloric acid was tested with indicator solution. Which box gives the colour produced?

d) The pH number of water is on the grid. Which box is it in?

Q7. N5
Which of the following oxides, when shaken with water, would leave the pH unchanged? You may wish to use the data booklet to help you.

A Carbon Dioxide
B Copper(II) oxide
C Sodium oxide
D Sulfur dioxide
4.2 Acid & Base Structures

**Acid Molecules**

If we look carefully at the *structures* of the substances that *dissolve in water* to produce *acidic solutions* we can see a pattern emerge.

*hydrogen chloride*

\[ \text{H}^{\delta+} \text{Cl}^{\delta-} \]

Firstly, they are all *covalent molecules* but all have a *very polar bond* involving a *hydrogen atom*.

*hydrogen carbonate*

\[ \text{O}^{\delta+} \text{H}^{\delta+} \text{C}^{\delta-} \text{O}^{-} \text{H}^{\delta+} \]

*hydrogen sulfate*

\[ \text{O}^{\delta+} \text{S}^{\delta+} \text{O}^{-} \text{H}^{\delta+} \]

*hydrogen nitrate*

\[ \text{O}^{\delta+} \text{N}^{\delta+} \text{O}^{-} \text{H}^{\delta+} \]

Secondly, most of these molecules have a *double covalent bond* from the *central atom* to an *oxygen atom* next to the *polar* \( \text{O}^{-} \text{H} \) bond.

\[ \text{H}^{\delta+} \text{O}^{\delta-} \]

Since *water molecules* are also *polar*, there will be *strong attractions* set up between the *water molecule* and the *acid molecule*.

As a result, the *acid molecule* will be *very soluble in water*.

However, whilst water *cannot* be *electrolysed* at *low voltages*, solutions of these *acid molecules* can be *electrolysed* and *always produce hydrogen gas at the negative electrode* (*cathode*).

\[ 2\text{H}^+_{(aq)} + 2e \rightarrow \text{H}_2(g) \]

This suggests that the following *change* has taken place:

\[ \text{covalent molecule} \rightarrow \text{ionic solution} \]

*acids are substances which dissolve in water to produce hydrogen ions, \( \text{H}^+_{(aq)} \).*
The strong attractions between water molecules and the acid molecules do much more than simply make them soluble in water. Water molecules are able to pull the acid molecules apart.

In the process, the shared electrons are completely transferred to the atom with the strong pull - in this case the chlorine atom. The acid is ionised - turned into ions dissolved in water.

Acids are substances which dissolve in water to produce hydrogen ions, \( H^+ \) (aq).

\[ \text{hydrogen chloride} \quad \text{H}^{\delta^+} \text{Cl}^{\delta^-} \]

\[ \text{hydrogen carbonate} \quad \delta^+ \text{H} - \text{O} - \text{C} - \text{O} - \text{H}^{\delta^+} \]

\[ \text{hydrogen sulfate} \quad \delta^+ \text{H} - \text{O} - \text{S} = \text{O} - \text{H}^{\delta^+} \]

\[ \text{hydrogen nitrate} \quad \text{O}^2- \text{N} - \text{O} - \text{H}^{\delta^+} \]

\[ \text{hydrochloric acid} \quad \text{Cl}^- \text{(aq)} \]

\[ \text{carbonic acid} \quad (\quad)_2 \text{CO}_3^{2-} \text{(aq)} \]

\[ \text{sulfuric acid} \quad (\quad)_2 \text{SO}_4^{2-} \text{(aq)} \]

\[ \text{nitric acid} \quad \text{H}^+ \text{NO}_3^- \text{(aq)} \]
Ammonia, or Amm, is the *only* base that starts off as a *cov mol* but when it *dis in wa* the *soul con* showing that *io* are formed. The pH of the *soul* is > showing that it has formed an *alk* solution.

The two *pol mol* att each other *strongly* enough to pull a *hydr ion* (H⁺) off the *wa* molecule.

The resulting *OH⁻ ion* makes the solution *alk*.

\[
\text{NH}_3 (g) + \text{H}_2\text{O} (l) \rightarrow \text{NH}_4^+ (aq) + \text{OH}^- (aq)
\]

A convenient way to make *amm* in the *lab* is to heat an ammonium *comp* with a *strong ba* or *alk*. *(The reverse of the reaction above)*

---

### Properties of Ammonia

- Ammonia is a *co less ga* at 20 °C
- Ammonia is the *only alk gas*
- Ammonia has a strong *pun smell*
- Ammonia is *le dense than air*
- Ammonia *dis in wa* to form an *alk solution* which will *con electricity*

As the ammonia gas cools and contracts, the water is slowly sucked up the tube. As the first drop appears at the end of the tube ALL the ammonia gas dissolves in this single drop of water.

As a result, a partial vacuum exists in the flask resulting in more water being sucked in very quickly to replace the ammonia gas molecules.

A 'fountain' is the result, showing that ammonia is extremely soluble in water. This is known as the *Fountain Experiment.*
More surprisingly, perhaps, is the fact that attraction between water molecules can also result in this covalent molecule being pulled apart to form ions.

The two polar molecules attract each other strongly enough to pull a hydrogen ion (H⁺) off one of the water molecules.

However an OH⁻ ion is also formed so water remains overall neutral, pH = 7.

Almost immediately, the ions will attract each other and the molecule will reform.

Water is a mixture of molecular and ionic, constantly breaking up and reforming.

The vast majority of water is molecular but there are always enough ions to make water a poor conductor.

acids are substances which dissolve in water to increase the concentration of hydrogen ions, H⁺(aq) - H⁺ concentration > 10⁻⁷

bases are substances which dissolve in water to increase the concentration of hydroxide ions, OH⁻(aq) - OH⁻ concentration > 10⁻⁷
**Acids & Bases**

### pH Numbers

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<th>pH</th>
<th>(10^{-7})</th>
<th>(\text{mol l}^{-1})</th>
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<td>(H^+)</td>
<td></td>
<td>(10^{-7})</td>
<td></td>
</tr>
<tr>
<td>(\text{mol l}^{-1})</td>
<td>(10^{-7})</td>
<td>(\text{OH}^-)</td>
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**Neutral solutions** have a **\(\text{pH} = 7\)**.

Neutral solutions have **equal amounts of** \(H^+\) ions and \(\text{OH}^-\) ions.

**Acid solutions** have a **\(\text{pH} < 7\)**.

Acid solutions have **more** \(H^+\) ions than \(\text{OH}^-\) ions, \(>\).

If you **dilute** an acid by adding **more water**, the pH will **increase** towards \(\text{pH} = 7\).

**Alkali solutions** have a **\(\text{pH} > 7\)**.

Alkali solutions have **more** \(H^+\) ions than \(\text{OH}^-\) ions, \(>\).

If you **dilute** an alkali by adding **more water**, the pH will **decrease** towards \(\text{pH} = 7\).

It would be tempting to assume that an acid of \(\text{pH} = 1\), eg stomach acid, was **three times** as **concentrated** as an acid of \(\text{pH} = 3\), eg lemon juice.

In fact, it is **\(10 \times 10^{-1}\)** i.e. **100 times** as concentrated.

*For each change in pH, a solution will become **ten times** more concentrated or **ten times** less concentrated.*

pH **pap** , **univ ind** and other **coloured** substances, such as **red** and **blue** paper and even the juice from **red cabbage**, can be used to measure the pH of solutions.

**Cond** devices, (\(H^+\) ions are good conductors), such as pH **met** and pH **probes** can also be used.
Q1.
a) What two features are found in most acid molecules?
_________________________________
_________________________________
b) Which of the following is an acid molecule?

A  H₂C₃O₂H
B  H₂C₃O₂⁻  CCl₂O₂⁻
C  H₂C₃O⁻  CCl₃O⁻
D  H₂C₃O２⁻  CCl₃O₂⁻

Q2.
The grid shows some statements which can be applied to different solutions.

A  It has a pH less than 7.
B  It conducts electricity.
C  It contains less OH⁻ ions than pure water.
D  It does not neutralise dilute hydrochloric acid.
E  When diluted the concentration of OH⁻ ions decreases.

Identify the two statements which are correct for an alkaline solution.

Q3.
What is the most likely pH value that would be obtained when zinc oxide is added to water?

(A You may wish to use page 5 of the data booklet to help you.)
A  5
B  7
C  9
D  11

Q4.
For each of the acids below
a) Write the formula for the molecule
b) Write the formulae for the two ions formed when it dissolves in water?

\[ \text{sulphurous acid} \]
\[ \text{molecular formula} \]
\[ \text{ions} \]
\[ \text{O=C-O-H} \]
\[ \text{nitrous acid} \]
\[ \text{molecular formula} \]
\[ \text{ions} \]
\[ \text{O=N-O-H} \]
\[ \text{chloric acid} \]
\[ \text{molecular formula} \]
\[ \text{ions} \]
\[ \text{OCl-O-H} \]

Q5.
A solution of 0·1 mol/l hydrochloric acid has a pH of 1.
a) What colour would universal indicator turn when added to a solution of hydrochloric acid?

b) Starting at pH 1, draw a line to show how the pH of this acid changes when diluted with water.

c) Calculate the number of moles of hydrochloric acid in 50 cm³ of 0·1 mol/l hydrochloric acid solution.
d) Magnesium carbonate can be used to neutralise acid:

\[ \text{MgCO}_3(s) + 2 \text{HCl} \rightarrow \text{MgCl}_2(aq) + \text{H}_2\text{O(l)} + \text{CO}_2(g) \]
i) Calculate the number of moles of MgCO₃ needed to neutralise 50 cm³ of 0·1 mol/l HCl solution.

ii) Calculate the mass of MgCO₃ needed to neutralise 50 cm³ of 0·1 mol/l HCl solution.
4.3 Reactions of Acids

With Metals

As you have probably learnt in earlier courses, reactive metals that are above hydrogen in the Reactivity Series are able to react with acids.

The gas produced burns with a squeaky noise. This shows that the gas is again hydrogen.

Reactive metals are able to force hydrogen ions to change back to hydrogen atoms. This allows the metal to take the place of the hydrogen and form a new substance called a Salt.

The sodium ion takes the place of the hydrogen ion to form the salt called sodium chloride.

Each acid has its own salts:

- hydrochloric acid, HCl → chlorides e.g. sodium chloride, NaCl
- sulfuric acid, H₂SO₄ → sulfates e.g. copper sulfate, CuSO₄
- nitric acid, HNO₃ → nitrates e.g. potassium nitrate, KNO₃

<table>
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<tr>
<th>Acid</th>
<th>+ metal</th>
<th>→ salt</th>
<th>+ hydrogen</th>
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<tbody>
<tr>
<td>e.g. magnesium + sulfuric acid</td>
<td>→</td>
<td>+ hydrogen</td>
<td></td>
</tr>
<tr>
<td>(s) + H₂SO₄(aq)</td>
<td>→</td>
<td>(aq) + (g)</td>
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</table>

We will learn more about this reaction if we firstly write an ionic equation and then remove spectator ions.

e.g. (s) + (H⁺)₂SO₄⁻(aq) → (aq) + (g)
| (s) + 2 H⁺(aq) | → | (aq) + (g) |
**With Metal Oxides**

In **me oxi** the **me** has already formed an **io** and will not **re** any more.

The **oxi ion**, O\(^{2-}\), reacts with the **hydr ions** in the **ac** to form water, H\(_2\)O.

This leaves the **me ion** to take the **hydr ions** place, so again a **sa** will be produced.

\[
\text{Acid} + \text{metal oxide} \rightarrow \text{salt} + \text{water}
\]

e.g. \[\text{iron (III)} + \text{nitric oxide} \rightarrow \text{iron(III)} + \text{water}\]

**balanced** \[\text{Fe}_2\text{O}_3 (s) + 6 \text{H}^+ (aq) \rightarrow 2 \text{Fe}^{3+} (aq) + 3 \text{H}_2\text{O} (l)\]

**ionic** \[(s) + 6 \text{H}^+ (aq) \rightarrow 2 \text{Fe}^{3+} (aq) + 3 \text{H}_2\text{O} (l)\]

**without spectator ions** \[(s) + 6 \text{H}^+ \rightarrow 3 \text{H}_2\text{O} (l)\]

**With Metal Carbonates**

In **me carb** the **me** has already formed an **io** and will not **re** any more.

The **carb ion**, CO\(_3^{2-}\), reacts with the **hydr ions** in the **ac** to form water and gas, CO\(_2\).

This leaves the **me ion** to take the hydrogens place, so again a **sa** will be produced.

\[
\text{Acid} + \text{metal carbonate} \rightarrow \text{salt} + \text{water} + \text{carbon dioxide}
\]

e.g. \[\text{calcium} + \text{hydrochloric carbonate} \rightarrow \text{calcium chloride} + \text{water} + \text{carbon dioxide}\]

**balanced** \[\text{CaCO}_3 (s) + 2 \text{HCl} (aq) \rightarrow \text{CaCl}_2 (aq) + \text{H}_2\text{O} (l) + \text{CO}_2 (g)\]

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**National 5**
Acids & Bases

ionic  

\[ \text{(s)} + (aq) \rightarrow (aq) + \text{H}_2\text{O (l)} + \text{CO}_2 (g) \]

without spectator ions

\[ \text{(s)} + \text{H}^+ (aq) \rightarrow \text{H}_2\text{O (l)} + \text{CO}_2 (g) \]

With Alkalis

Alkalis are solutions which contain hydroxide ions.

Most alkalis are made by dissolving a metal oxide in water - though only those in Group 1 are very soluble.

\[
\begin{align*}
\text{Na}_2\text{O (s)} & + \text{H}_2\text{O (l)} \rightarrow \text{NaOH (aq)} \\
\text{K}_2\text{O (s)} & + \text{H}_2\text{O (l)} \rightarrow \text{(aq)} \\
\text{CaO (s)} & + \text{H}_2\text{O (l)} \rightarrow \text{(aq)} \\
\text{MgO (s)} & + \text{H}_2\text{O (l)} \rightarrow \text{(aq)}
\end{align*}
\]

soda      sodium hydroxide
pearl ash     potassium hydroxide
calcium hydroxide
magnesia              magnesium hydroxide

It is the hydroxide ion which will react with the acid, and water is the product

<table>
<thead>
<tr>
<th>Acid</th>
<th>+</th>
<th>alkali</th>
<th>→</th>
<th>salt</th>
<th>+</th>
<th>water</th>
</tr>
</thead>
<tbody>
<tr>
<td>e.g</td>
<td>sodium hydroxide</td>
<td>+</td>
<td>sulfuric acid</td>
<td>→</td>
<td>+</td>
<td>water</td>
</tr>
<tr>
<td></td>
<td>(aq)</td>
<td>+</td>
<td>H$_2$SO$_4$ (aq)</td>
<td>→</td>
<td>(aq)</td>
<td>+</td>
</tr>
<tr>
<td>ionic</td>
<td>(aq)</td>
<td>+</td>
<td>(H$^+$)$_2$ SO$_4^{2-}$ (aq)</td>
<td>→</td>
<td>(aq)</td>
<td>+</td>
</tr>
</tbody>
</table>

Once again, the metal ion to take the hydrogens place, so again a salt will be produced. The metal ions are spectator ions as are the sulfate ions in the example above.

without spectator ions

\[ (aq) + \text{H}^+ (aq) \rightarrow \text{H}_2\text{O (l)} \]
Making Salts

Salt Preparation

Because of the large number of, in particular, metal oxides and carbonates that it is possible to react easily with a number of acids, a whole range of ‘new substances’ can be made using acid reactions. If we include, precipitation reactions (met earlier in the course) there are very few compounds that cannot be made quickly and easily.

This is basically a Problem Solving activity that will test your knowledge of acid reactions, your use of solubility tables, your appreciation of practical considerations (such as ensuring complete reaction) and your knowledge of separation techniques.

There are 3 parts to salt preparation
1) Choice of Reaction
2) Reaction Method
   and 3) Separation of Salt produced

1) Possible Reactions
   a) Acid + (solid) Metal
   b) Acid + (solid) Oxide, Hydroxide or Carbonate
   c) Acid + Alkali (solution)
   d) Precipitation (solutions, one of which may be acid)

2) Reaction Methods & 3) Separation of Salt

Solid Metals, Oxides, Hydroxides & Carbonates

The solid is added spatula by spatula, stirring all the time, until there is an obvious layer of unreacted (excess) solid lying at the bottom. The acid may need to be heated to speed up the reaction.

The excess solid must now be separated from the salt solution by filtering. The solid trapped in the filter paper can be discarded.

The salt solution can now be heated until all the water has evaporated away leaving solid salt powder. If preferred, the solution can be left to evaporate slowly in which case salt crystals will form.

Alkali solutions

The acid will need to be added slowly and carefully (eventually drop by drop) until the indicator just changes colour.

Then, to another flask, the exact same volumes will be reacted but without the indicator present.

As above.
Using the Data Book, two suitable solutions will need to be made up. Each solution will provide one half of the salt to be made. Once made, the two solutions are simply mixed together.

Sometimes a choice of methods is available. To help choose the ‘best’ method the following summary may be useful.

<table>
<thead>
<tr>
<th>Reaction</th>
<th>'Advantages'</th>
<th>'Disadvantages'</th>
<th>'Suitability'</th>
</tr>
</thead>
<tbody>
<tr>
<td>solid metals</td>
<td>Easy to ensure 'complete' reaction - excess metal left over at end.</td>
<td>Must filter excess metal.</td>
<td>Not suitable if salt is insoluble - too difficult to separate from excess solid metal.</td>
</tr>
<tr>
<td>solid oxides, hydroxides, carbonates</td>
<td>Easy to ensure 'complete' reaction - excess solid left over at end.</td>
<td>Must filter excess solids.</td>
<td>Not suitable if salt is insoluble - too difficult to separate from excess solid.</td>
</tr>
<tr>
<td>alkali solutions</td>
<td>Reaction immediate. No need to filter excess solids.</td>
<td>Difficult to ensure 'exact' neutralisation. Technique may take a very long time.</td>
<td>Very limited choice of alcalis - so limited number of salts can be prepared by this method.</td>
</tr>
<tr>
<td>precipitation from solutions</td>
<td>Reaction extremely quick.</td>
<td>None really.</td>
<td>Limited to insoluble salts only.</td>
</tr>
</tbody>
</table>

**Examples**

**To prepare copper sulfate**
1. The Data Book will tell you that copper sulfate is soluble, so precipitation is out.
2. Acid to use: - sulfuric acid
3. Copper metal - no reaction with acid.
4. Copper oxide/hydroxide is insoluble so there is no alkali solution, so titration is out.
5. Best method would be to add solid copper oxide/hydroxide/carbonate to sulfuric acid.

**To prepare zinc chloride**
1. The Data Book will tell you that zinc chloride is soluble, so precipitation is out.
2. Acid to use: - hydrochloric acid
3. Zinc reacts slowly with acid.
4. Zinc oxide/hydroxide is insoluble so there is no alkali solution, so titration is out.

**To prepare silver chloride**
1. The Data Book will tell you that silver chloride is insoluble, so precipitation is the best method.
2. Use the Data Book to find a soluble silver compound eg silver nitrate.

Use the Data Book to find a soluble chloride compound eg sodium chloride.
Q1. N5

AgNO₃(aq) + KCl(aq) → AgCl(s) + KNO₃(aq)

Which of the following are the spectator ions in this reaction?

A  Ag⁺ and Cl⁻
B  K⁺ and NO₃⁻
C  Ag⁺ and NO₃⁻
D  K⁺ and Cl⁻

Q2. SC

Copy and complete the following equations

lithium + hydrochloric → + water
hydroxide acid

aluminium + nitric → + water
oxide acid

strontium + nitric → + water +
carbonate acid

Q3. SC

Copy, complete and balance the following equations

Ca(OH)₂ + HCl → + H₂O

CuO + → Cu(NO₃)₂ +

+ H₂SO₄ → K₂SO₄ + + CO₂

Q4. SC

Lead(II) nitrate solution reacts with potassium iodide solution to give a yellow solid.

Pb²⁺(aq) + 2NO₃⁻(aq) + 2K⁺(aq) + 2I⁻(aq) → PbI₂(s) + 2K⁺(aq) + 2NO₃⁻(aq)

Identify the two spectator ions in the reaction.

_____________ and ______________

Q5. N5

Which of the following salts can be prepared by a precipitation reaction?

You may wish to use the data booklet to help you.

A  Barium sulfate
B  Lithium nitrate
C  Calcium chloride
D  Ammonium phosphate

Q6. Int2

Copy and complete the following equations, clearly showing which of the products is the precipitate.

BaCl₂(aq) + K₂SO₄(aq) → +

CuSO₄(aq) + Na₂CO₃(aq) → +

AgNO₃(aq) + KCl(aq) → +

Q7. Int2

Reactions can be represented using ionic equations. Which ionic equation shows a neutralisation reaction?

A  2H₂O(l) + O₂(g) + 4e → 4OH⁻(aq)
B  H⁺(aq) + OH⁻(aq) → H₂O(l)
C  SO₂(g) + H₂O(l) → 2H⁺(aq) + SO₄²⁻(aq)
D  NH₄⁺(s) + OH⁻(s) → NH₃(g) + NH₄⁺(s)

Q8. A solution of sulphuric acid can be used to neutralise a solution of sodium hydroxide.

a) What is the pH of the solution when it is exactly neutral. __________

b) What is the name of the salt formed in the neutralisation reaction? ______________

c) Balance the following equation for the reaction.

(H⁺)₂SO₄²⁻ + NaOH → (Na⁺)₂SO₄²⁻ + H₂O

d) Rewrite the equation, omitting the spectator ions.

+ →
4.4 Quantitative Analysis

Titrations

It is quite common for a Chemist to be asked to measure how much acid or alkali is present in a solution - for example, how much acid is present in lemonade.

A technique called titration is used, and you may be expected to demonstrate your ability to carry out a titration.

A carefully measured volume of the lemonade would be placed in the flask. An indicator that will change colour is also added.

An alkali whose concentration is accurately known would be placed in a burette, and then added to the flask.

Eventually the alkali would be added drop by drop until the indicator changes colour to show that the acid in the lemonade has been neutralised.

Before the alkali can be used to determine how much acid is present in the lemonade, it must be standardised - titrated against a Standard Solution of a suitable acid, such as potassium hydrogen phthalate (KHP).

An unusual acid like this is chosen because it is extremely stable, very soluble and can be made to a very high level of purity.

An analytical balance is used to very accurately weigh out a calculated mass of the chemical.

From this mass, the number of moles of chemical can be calculated and used to calculate the concentration of the solution made.

Molecular Formula:

**H**<sub>2</sub>**C**<sub>4</sub>**H**<sub>4</sub>**O**<sub>4</sub>K<sup>+</sup>

**H**<sub>2</sub>**C**<sub>4</sub>**H**<sub>4</sub>**O**

**OH**

**O**

potassium hydrogen phthalate
mass of KHP = 10.00 g

\[
g \quad \rightarrow \quad 1 \text{ mole}
\]

\[
10.00 \text{ g} \quad \rightarrow \quad 1 \times 10.00
\]

= moles

vol of flask = 500 ml = 0.5 l

\[
C = \frac{n}{V}
\]

= / 0.5

= mol l\(^{-1}\)

Technique - Standardisation of alkali (NaOH)

1. Filling the burette

   - Set up the burette carefully in a stand. Clamp gently!

   - Collect a beakerful of the alkali you are going to use and another empty waste beaker.

   - Pour just a little of the alkali into the burette to rinse it. Pour it out into your waste beaker.

   - With the valve closed, fill up the burette to just above the zero line. Then remove the funnel.

   - Open the valve slightly and let the alkali drip into your waste beaker until the bottom of the curved surface is on or below the zero line.

6. Read your starting volume. Read with your eye level with the curved surface. Make a note of this reading. The burette is now ready for use.
2. **Using the pipette**  

   The pip is used to accurately measure out the same vol of KHP every time.

1. Collect a bea ful of the KHP you are going to use, and a conical fla .
2. Use the fil to suck the KHP above the fill mark.
3. Holding the pip above the beaker, slowly let the KHP drip out until the bot of the cur surface is on the fill mark.
4. Carefully transfer the KHP in the pip into your flask.

   A tiny amount will remain inside the tip. This is supposed to happen.

   Dip the tip of the pip into the KHP and some more will come out. Any still left in the pipette is allowed for.
5. Add a few drops of indi .

3. **Doing the Titration**  

   The aim is to find out exactly what vol of alkali is needed to neut the known vol of KHP.

1. Put a piece of wh paper under the fla . It will help you to see the col of indi .
2. Start by adding the alkali 5 ml at a time. You should see the indi col change but then return quickly.
3. If the col takes longer to return, add less alkali next time. Ideally you should add one drop of alkali and see the indi change permanently.
4. Write down the final burette reading to at least the nearest 0.1 ml.

Remember that a burette reads downwards.

The burette opposite is reading ml (and not ml).

5. You will now need to repeat your titration with a freshly pipetted sample of KHP.

Knowing the answer from the first attempt should allow you to quickly add enough alkali to nearly neutralise the potassium hydrogen phthalate. Then you can add more alkali drop by drop to get an accurate result.

6. You should record your results in a table similar to the one on the right.

Your first attempt is often a ‘rough’ titration as you will often add too much alkali at a time.

Later attempts should produce results to the nearest drop.

You have to continue repeating the titration until you get two results at least within 0.1 ml of each other.

7. You should finish off by quoting your conclusion in terms of:

“It takes 21.7 ml of NaOH to neutralise 20 ml of KHP”.

If you have two, or more, answers close to each other then it may be better to use their average as your final answer.

To get similar results each time, (the aim of this technique), you will need to work hard to ensure that you pipette exactly the same volume of potassium hydrogen phthalate each time.

<table>
<thead>
<tr>
<th>Attempt number</th>
<th>Starting volume (ml)</th>
<th>Final volume (ml)</th>
<th>Volume added (ml)</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>0.1</td>
<td>22.0</td>
<td>21.9</td>
</tr>
<tr>
<td>2</td>
<td>22.0</td>
<td>43.5</td>
<td>21.5</td>
</tr>
<tr>
<td>3</td>
<td>0.2</td>
<td>21.9</td>
<td>21.7</td>
</tr>
<tr>
<td>4</td>
<td>21.9</td>
<td>42.6</td>
<td>21.7</td>
</tr>
</tbody>
</table>
Q1.

The table below shows the colours of various indicators at different pH values.

<table>
<thead>
<tr>
<th>indicator</th>
<th>pH 1</th>
<th>colour 1</th>
<th>pH 2</th>
<th>colour 2</th>
</tr>
</thead>
<tbody>
<tr>
<td>bromophenol blue</td>
<td>3</td>
<td>yellow</td>
<td>4.5</td>
<td>blue</td>
</tr>
<tr>
<td>phenolphthalein</td>
<td>8</td>
<td>colourless</td>
<td>10</td>
<td>pink</td>
</tr>
<tr>
<td>methyl orange</td>
<td>3</td>
<td>red</td>
<td>4.5</td>
<td>yellow</td>
</tr>
<tr>
<td>thymol blue</td>
<td>6</td>
<td>yellow</td>
<td>7.5</td>
<td>blue</td>
</tr>
</tbody>
</table>

The table below shows the pH of some solutions.

<table>
<thead>
<tr>
<th>solution</th>
<th>pH</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.1 M hydrochloric acid</td>
<td>1.0</td>
</tr>
<tr>
<td>0.1 M ethanoic acid</td>
<td>5.0</td>
</tr>
<tr>
<td>0.1 M ammonia</td>
<td>10.0</td>
</tr>
<tr>
<td>0.1 M sodium hydroxide</td>
<td>12.5</td>
</tr>
</tbody>
</table>

a) Complete the table below to show the colours of the indicators in the solutions.

<table>
<thead>
<tr>
<th>indicator</th>
<th>solution</th>
<th>colour</th>
</tr>
</thead>
<tbody>
<tr>
<td>bromophenol blue</td>
<td>0.1 M hydrochloric acid</td>
<td></td>
</tr>
<tr>
<td>phenolphthalein</td>
<td>0.1 M ethanoic acid</td>
<td></td>
</tr>
<tr>
<td>methyl orange</td>
<td>0.1 M ammonia</td>
<td></td>
</tr>
<tr>
<td>thymol blue</td>
<td>0.1 M sodium hydroxide</td>
<td></td>
</tr>
</tbody>
</table>

b) Name one indicator which turns the same colour in both ethanoic acid and sodium hydroxide.

__________________

c) Which two indicators turn the same colour in hydrochloric acid

__________________

Q2.

One of the solids often used in Antacid Tablets to treat indigestion is magnesium hydroxide.

A pupil decided to find out how much of the solid would be needed to neutralise some acid.

a) Complete the equation for the reaction of magnesium hydroxide with hydrochloric acid

\[
\text{Mg(OH)}_2 (aq) + 2\text{HCl} (aq) \rightarrow \quad + \quad \text{H}_2\text{O} (l)
\]

b) Calculate the number of moles of HCl present.

c) Calculate the number of moles of Mg(OH)_2 needed.

d) Calculate the mass of Mg(OH)_2 needed.

25 cm³ of 2M HCl
Knowledge Met in this Topic

Common household acids and bases
- **Acids:** vinegar, citrus fruits, cola drinks etc
- **Bases:** lime, oven cleaner, bleach, bicarbonate of soda, soap, ammonia

**Water**
- Water has *polar covalent* molecules which can break up (*dissociate*) to produce $H^+_{(aq)}$ and $OH^-_{(aq)}$ ions as shown:
  \[ H_2O \ (l) \rightleftharpoons H^+_{(aq)} + OH^-_{(aq)} \]
  where the symbol $\rightleftharpoons$ shows that the reaction occurs in both directions, it is *reversible*.
- At any time only a very few water molecules are *dissociated* into free ions, and
  \[ \text{number of } H^+_{(aq)} \text{ ions} = \text{number of } OH^-_{(aq)} \text{ ions} \]  
  (neutral)

**Important acids**
- Most acids start off as *polar covalent molecules* which *dissociate in water* to produce *extra hydrogen ions*, $H^+_{(aq)}$.
- *Hydrochloric* acid: $HCl_{(g)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$
- *Sulphuric* acid: $H_2SO_4_{(l)} \rightarrow 2H^+_{(aq)} + SO_4^{2-}_{(aq)}$
- *Nitric* acid: $HNO_3_{(l)} \rightarrow H^+_{(aq)} + NO_3^-_{(aq)}$
- As a result:
  \[ \text{number of } H^+_{(aq)} \text{ ions} > \text{number of } OH^-_{(aq)} \text{ ions} \]  
  (acidic)

**Acidic Oxides**
- Oxides of *non-metals* which *dissolve* produce *acidic* solutions, $H^+_{(aq)} > OH^-_{(aq)}$
e.g., $CO_2$, $SO_2$ and $NO_2$.
- *Non-metal* oxides are the main cause of *acid rain*.

**Basic Oxides**
- Oxides of *metals* which *dissolve* react with water to produce *hydroxides*
  \[ \text{metal oxide} + \text{water} \rightarrow \text{metal hydroxide} \]
- Oxides of *metals* which *dissolve* produce *alkaline* solutions, $OH^-_{(aq)} > H^+_{(aq)}$
- *All* the oxides of *Group 1* metals are very soluble, only *some* from *Group 2* are soluble.
**The pH scale**

- pH scale is a number scale that shows how acidic or alkaline a solution is and runs from below 0 to above 14.
- *Universal indicator, pH paper or a pH meter* can show the pH of a solution.
- **Acids** pH less than 7, pH < 7  
  **Neutral** pH equals 7, pH = 7  
  **Alkalis** pH more than 7, pH > 7
- When *acids* dissolve in water they produce *extra hydrogen ions*, $\text{H}^+_{(aq)}$  
- When *bases* dissolve in water they produce *extra hydroxide ions*, $\text{OH}^-_{(aq)}$  
- Pure water and all *neutral* solutions contain a tiny but *equal* concentration of hydrogen and hydroxide ions, $\text{H}^+_{(aq)} = \text{OH}^-_{(aq)}$  
- An acid *solution* contains *more hydrogen ions* than hydroxide, $\text{H}^+_{(aq)} > \text{OH}^-_{(aq)}$  
- An alkali *solution* contains *less hydrogen ions* than hydroxide, $\text{H}^+_{(aq)} < \text{OH}^-_{(aq)}$  
- *Diluting* acids or alkalis will reduce the concentration of either $\text{H}^+_{(aq)}$ or $\text{OH}^-_{(aq)}$ ions and *move the pH towards 7*.

**Neutralisation**

- *Neutralisation* is a reaction in which a *base* reacts with an *acid* to *form water*. A *salt* is also formed in this reaction.  
- *Neutralisation* is a *reaction* in which the pH of a solution moves towards 7.  
- **Neutralisers** which can be used to *neutralise* an acid include the following *bases* (insoluble solids) and *alkalis* (soluble solutions):
  - solid metal oxides (*base*),  
    - metal oxide + an acid $\rightarrow$ salt + water  
  - soluble metal oxides / metal hydroxide *solutions* (*alkali*),  
    - metal hydroxide + an acid $\rightarrow$ salt + water  
  - soluble metal carbonates (*alkali*) / insoluble metal carbonates (*base*),  
    - metal carbonate + an acid $\rightarrow$ salt + water + $\text{CO}_2$

**Everyday examples of neutralisation**

- *Lime* (calcium oxide) is used to *reduce acidity* in soil and water.  
- Cures for *acid indigestion* contain neutralisers such as *calcium carbonate*.  
- *Wasp stings* (base) can be neutralised using *vinegar* (acid).  
- *Bee stings* and *Ant Bites* (acid) can be neutralised using *baking soda* (base).
**Acids & Bases Topic 4**

**Bases and alkalis**

- A **base** is a substance that neutralises an acid.
- An **alkali** is a base that dissolves in water.

**Salts**

- **Salts** are ionic compounds formed in reactions between acids and bases.
- A metal ion will have replaced the hydrogen in an acid.
- Hydrochloric acid $\text{HCl}$ forms **chlorides** e.g. $\text{NaCl}$
- Sulphuric acid $\text{H}_2\text{SO}_4$ forms **sulphates** e.g. $\text{CuSO}_4$
- Nitric acid $\text{HNO}_3$ forms **nitrates** e.g. $\text{KNO}_3$
- Reacting acids is a good way of making salts.
- **Soluble Salts** can be made by reacting excess insoluble metal oxides or metal carbonates with a suitable acid. e.g.

\[
2 \text{H}^+ \text{Cl}^-(aq) + \text{Cu}^{2+} \text{O}^{2-}(s) \rightarrow \text{Cu}^{2+}(\text{Cl}^-)_{2(aq)} + \text{H}_2\text{O}(l)
\]

- At the end of the reaction the mixture is filtered and the soluble salt can be separated by evaporation to dryness.
- **Insoluble Salts** can be made by precipitation - two solutions combine to produce an insoluble precipitate of the salt. e.g.

\[
2 \text{H}^+ \text{Cl}^-(aq) + \text{Ba}^{2+}(\text{OH}^-)_{2(aq)} \rightarrow \text{Ba}^{2+}(\text{Cl}^-)_{2(s)} + 2 \text{H}_2\text{O}(l)
\]

- **Insoluble Salts** can also be made by precipitation reactions that do not involve an acid or an alkali.

\[
2 \text{K}^+ \text{I}^-(aq) + \text{Pb}^{2+}(\text{NO}_3^-)_{2(aq)} \rightarrow \text{Pb}^{2+}(\text{I}^-)_{2(s)} + 2 \text{K}^+\text{NO}_3^-(aq)
\]

- At the end of the reaction the precipitate of the insoluble salt can be separated by filtration and then dried.

**Neutralisation Reactions**

- **Neutralisation reactions** involve spectator ions that remain unchanged by the reaction.
- **Reaction equations** can be used to identify spectator ions.

<table>
<thead>
<tr>
<th>Reactions</th>
<th>Products</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Acid</strong> + <strong>Alkali</strong> → <strong>Water</strong> + <strong>‘salt’</strong></td>
<td>$\text{H}^+(aq) + \text{OH}^-(aq) \rightarrow \text{H}_2\text{O}$</td>
</tr>
<tr>
<td><strong>Acid</strong> + <strong>Oxide/Hydroxide</strong> → <strong>Water</strong> + <strong>‘salt’</strong></td>
<td>$2\text{H}^+(aq) + \text{O}^{2-}(aq) \rightarrow \text{H}_2\text{O}$</td>
</tr>
</tbody>
</table>
Acids & Bases

• Acid + Carbonate → Water + Carbon dioxide + ‘salt’
  
  eg  \[ 2H^+ \text{Cl}^- (aq) + Ca^{2+} CO_3^{2-} (s) \rightarrow H_2O (l) + CO_2 (g) + Ca^{2+} Cl^- (aq) \]

removing spectator ions:-  
\[ 2H^+ (aq) + CO_3^{2-} (s) \rightarrow H_2O (l) + CO_2 (g) \]

• Acid rain reacts with carbonate rocks such as marble (statues) and limestone.

Ammonia (a special base)

• Ammonia is a colourless gas with a sharp, unpleasant (pungent) smell.
• Ammonia is a very soluble in water producing an alkaline solution.
  
  \[ NH_3 + H_2O \rightarrow NH_4^+ + OH^- \]

• Ammonia is the only common alkaline gas.
• Ammonia can neutralise an acid and form an ammonium salt.
  
  \[ NH_3 + H_2SO_4 \rightarrow (NH_4)_2SO_4 \text{ (ammonium sulphate)} \]

• Ammonia gas can be produced when an ammonium salt is heated with an alkali e.g.  
  \[ \text{NH}_4\text{Cl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O} + \text{NH}_3 \]

Titrations

• Titration is an analytical technique in which volumes of chemicals involved in a chemical reaction can be determined accurately.
• Instruments such as burettes and pipettes are used to determine accurate volumes.
• The concentration of one of the chemicals must be known. This is done by accurate weighing and careful dissolving in a volumetric flask.
• A solution whose concentration is accurately known is a Standard Solution.
• A Standard Solution can be used to determine the concentration of another chemical if we know how they react - know the balanced equation.
• An indicator is used to show the end-point of the reaction.
• In titrations involving acids and alkalis, the end-point is when the solutions have neutralised each other.
• Titration allows the concentration of one reactant to be calculated as long as the concentration of the other reactant is known.
• Titration can be used to produce a soluble salt. Once the volumes of acid and alkali needed are know, the titration can be repeated without indicator.
• The soluble salt produced can be separated by evaporation to dryness.
Calculations (separate Calculations booklet)

- Knowing the formula for a substance, the **Formula Mass** can be calculated using *relative atomic masses* using values contained in the Data Booklet.

- The mass of **1 mole** of a substance is equal to the **Formula Mass** expressed in **grammes** - the **gramme formula mass (gfm)**.

- The **mass** of any **number of moles** of a chemical can be calculated:

\[
\text{mass} = \text{no. of moles} \times \text{gfm}
\]

- The **number of moles** in any **mass** of a chemical can be calculated:

\[
\text{no. of moles} = \frac{\text{mass}}{\text{gfm}}
\]

- Calculations involving **solutions** must have volumes expressed in **litres**.

\[
\text{concentration} = \frac{\text{no. of moles}}{\text{volume}}
\]

\[
\text{no. of moles} = \text{concentration} \times \text{volume}
\]

- **Titration** calculations can be done in steps or by using formulae such as:

\[
\frac{C_{\text{ACID}} \times V_{\text{ACID}}}{n_{\text{ACID}}} = \frac{C_{\text{ALK}} \times V_{\text{ALK}}}{n_{\text{ALK}}}
\]

where \(n\) = **number of moles in the balanced equation**
Q1.

Below is information about six chemicals.

<table>
<thead>
<tr>
<th>chemical</th>
<th>state at 20 °C</th>
<th>pH in water</th>
<th>reaction with water</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td>gas</td>
<td>1</td>
<td>none</td>
</tr>
<tr>
<td>B</td>
<td>liquid</td>
<td>7</td>
<td>none</td>
</tr>
<tr>
<td>C</td>
<td>solid</td>
<td>4</td>
<td>none</td>
</tr>
<tr>
<td>D</td>
<td>solid</td>
<td>8</td>
<td>forms salt, carbon dioxide and water</td>
</tr>
<tr>
<td>E</td>
<td>solid</td>
<td>14</td>
<td>forms a salt and water</td>
</tr>
<tr>
<td>F</td>
<td>solid</td>
<td>no reaction</td>
<td>fizzes</td>
</tr>
</tbody>
</table>

Use the table to write the letter of the chemical substance which:

- $a)$ forms the most strongly acidic solution
- $b)$ forms a neutral solution
- $c)$ is a metal
- $d)$ forms a solution which turns pH paper orange

Q2.

Equations are used to represent chemical reactions.

A) Zn(s) $\rightarrow$ Zn$^{2+}$(aq) + 2e$^-$
B) $C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l)$
C) $SO_2(g) + H_2O(l) \rightarrow 2H^+(aq) + SO_3^{2-}(aq)$
D) $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$
E) $SO_2^{2-}(aq) + 2H^+(aq) + 2e^- \rightarrow SO_3^{2-}(aq) + H_2O(l)$

- $a)$ Identify the equation which represents the formation of acid rain.
- $b)$ Identify the equation which involves sulphuric acid.

Q3.

Reactions can be represented using ionic equations. Which ionic equation shows a neutralisation reaction?

- $A)$ $2H_2O(l) + O_2(g) + 4e^- \rightarrow 4OH^-(aq)$
- $B)$ $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$
- $C)$ $SO_3^{2-}(aq) + H_2O(l) \rightarrow 2H^+(aq) + SO_3^{2-}(aq)$
- $D)$ $NH_4^+(aq) + OH^-(aq) \rightarrow NH_3(g) + NH_4^+(aq)$

Q4.

The grid shows some ions.

A) Al$^{3+}$  B) Cl$^-$  C) Li$^+$  D) H$^+$  E) Br$^-$  F) OH$^-$

- $a)$ Identify the two ions which combine to form an insoluble compound.
- $b)$ Identify the ion present in all alkaline solutions.

Q5.

A solution of accurately known concentration is more commonly known as a

- A) correct solution
- B) precise solution
- C) standard solution
- D) prepared solution.
CONSOLIDATION QUESTIONS       B

Q1 Both ammonia molecules and hydrogen chloride molecules are described as being polar.

a) What is meant by the word polar, as used in this context.

b) Complete the formula for hydrogen chloride to show its polar characteristics.

\[ \text{H} \quad \text{—} \quad \text{Cl} \]

c) Ammonia gas \(\text{NH}_3(\text{g})\), can be dissolved in water to form concentrated ammonia solution. Hydrogen chloride gas \(\text{HCl}(\text{g})\), can be dissolved in water to form concentrated hydrochloric acid.

If both bottles are placed next to each other in a fume cupboard and the stoppers removed, both liquids evaporate and a white cloud is formed where the two gases meet.

i) State the colour of the pH paper at X and Y.

pH paper X   ____________        pH paper Y   ____________

ii) The white cloud appears because the gases react to form a salt. Name the salt formed.

___________________________________________
**CONSOLIDATION QUESTIONS**

**Q1** Sulfur dioxide is an important industrial chemical.

Sulfur dioxide dissolves in water to produce sulfurous acid.

\[
\text{SO}_2(g) + \text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{SO}_3(aq)
\]

\text{a)} Explain the change in the pH of the solution as sulfur dioxide dissolves.

\text{b)} The graph shows the solubility of sulfur dioxide at different temperatures.

Describe the general trend in solubility as the temperature of the water increases.

**Q2** A student was given two solutions of sodium carbonate, one solution with a concentration of 0·1 mol l\(^{-1}\) and the other with a concentration of 0·2 mol l\(^{-1}\).

*Using your knowledge of chemistry*, suggest how the student could distinguish between the solutions.

**Q3** Carbonated water, also known as sparkling water, is water into which carbon dioxide gas has been dissolved. This process is called carbonating.

A group of students are given two brands of carbonated water and asked to determine which brand contains more dissolved carbon dioxide.

*Using your knowledge of chemistry*, describe how the students could determine which brand of carbonated water contains more dissolved carbon dioxide.

**Q4** A student reacted aluminium with dilute nitric acid.

\[
2\text{Al}(s) + 6\text{HNO}_3(aq) \rightarrow 2\text{Al(NO}_3)_3(aq) + 3\text{H}_2(g)
\]

\text{i)} Circle the formula for the salt in the above equation.

\text{ii)} 1 mole of hydrogen gas has a volume of 24 litres.

Calculate the volume of hydrogen gas, in litres, produced when 0·01 moles of aluminium react with dilute nitric acid.
CONSOLIDATION QUESTIONS

Q1  A student investigated the reaction between dilute sulphuric acid and sodium carbonate.

His experiment involved determining the concentration of sodium carbonate solution by titration.

The results showed that 20 cm$^3$ of sulphuric acid was required to neutralise the sodium carbonate solution.

a) Calculate the number of moles of sulphuric acid in this volume.

________________ mol

b) One mole of sulphuric acid reacts with one mole of sodium carbonate.

Using your answer from part a), calculate the concentration, in mol/l, of the sodium carbonate solution.

________________ mol/l

c) Name the salt produced when dilute sulphuric acid reacts with sodium carbonate.

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