NAME: ______________________________

TALENT 100
HSC Success. Simplified.

Preliminary Chemistry.

CHEMISTRY SUCCESS. SIMPLIFIED.

Reactive Chemistry III

Theory booklet

Visit my-mentor.com.au for all your online learning resources and support

1300 999 100 | talent-100.com.au | Chatswood | Epping | Hurstville | Burwood

SYLLABUS OUTCOMES

Inquiry question: How is the reactivity of various metals predicted?

By the end of this lesson, students should be able to:

- ☐ Explain the exothermic or endothermic nature of a given reaction by considering law of conservation of energy and understanding how breaking and forming bonds causes heat energy to be absorbed and released. (adapted from ACSCH037)
- ☐ Conduct practical investigations to compare the reactivity of a variety of metals in:
  - ☐ water
  - ☐ dilute acid (ACSCH032, ACSCH037)
  - ☐ oxygen
  - ☐ other metal ions in solution
- ☐ Construct a metal activity series using the data obtained from practical investigations and compare this series with that obtained from standard secondary-sourced information (ACSCH103)

Name: ____________________ Class time: ______________ Mentor: ___________________

__
6

QUIZ

Question 1 (1 mark)

Write the reaction when sulfuric acid reacts with calcium carbonate.

………

………

Question 2 (2 marks)

Identify the transitional pH ranges and colours for the indicators litmus, phenolphthalein, bromothymol blue and methyl orange.

………

………

………

………

………

………

………

………

Question 3 (3 marks)

2.8 grams of solid potassium hydroxide is dissolved in 200mL of water. The solution is then mixed with a solution of hydrochloric acid of volume 300mL and pH 0.78. What is the resultant pH of the solution?

………

………

………

………

………

………

………

………

………

………

………

………

GENERAL PRINCIPLES OF REACTIVITY

MAKING AND BREAKING BONDS

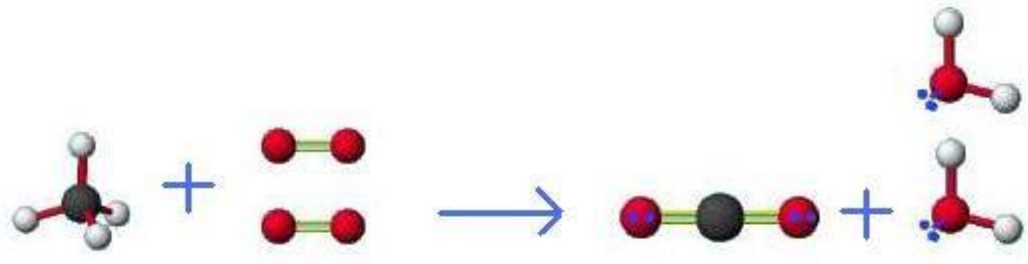
On the macroscopic level, all chemical reactions can be broken down into two fundamental processes:

1. The chemical bonds in reactant molecules **are broken** when there is sufficient energy in the system (i.e. heat energy) to initiate the reaction.
2. The atoms are rearranged to **form new** chemical bonds and the products.

Consider the complete combustion of methane represented in two forms. The balanced chemical equation for the reaction is:

$$CH_{4\,(g)} + 2O_{2\,(g)} \rightarrow CO_{2\,(g)} + 2H_2O_{\,(g)}$$

In the context of this discussion, the ball and stick model is a useful tool we can use to explain the bond breaking and making processes. (*Note: Black is Carbon, White is Hydrogen, and Red is Oxygen*)



With a sufficient energy input into the system, the C-H bonds in methane and the O=O double bonds oxygen gas are broken. As the reaction proceeds, new chemical bonds are formed, specifically C=O double bonds and O-H bonds. It is clear that the *same* atoms are used to form the products, carbon dioxide and water.

In the process described above there is a *change in the energy* of the system. These changes may be positive or negative. So why does a combustion reaction release energy?

BONDS AND ENERGY

In chemical reactions, most energy transfers are related to the process of breaking and forming new chemical bonds.

Breaking bonds ………………………………………………… whereas forming bonds ………………………… …………………………

Think of the above expression as an accounting system to follow all processes that bring about a change in energy. Importantly, there are two possible outcomes when describing the *overall* energy change in a chemical system:

In **exothermic** reactions – ones which …………………………………………………… **from the** …………………………………………………… **into the** ……………………….. – the amount of energy released by bond formation **is greater** than the amount of energy required to break the bonds of reactant molecules.

In **endothermic** reactions – ones which …………………………………………………… **from the** ……………………….. **into the** …………………………………………………… – the energy required to break the bonds **exceeds** the amount of energy that is released when the products form.

In light of the question posed earlier, it follows that the **energy released** when bonds are formed in a combustion reaction **exceeds** the energy required to break the bonds of reactant molecules. We can therefore class the process of combustion as an exothermic reaction.

These concepts will be covered in greater detail in the next module as they are used to predict the spontaneity of chemical reactions. Over the next three lessons, we will study the reactivity of metals and determine the energy transfers that occur during a range of chemical reactions.

THE REACTIVITY OF METALS

If you were to drop 100 g of pure sodium metal into a lake, you would witness a loud explosion as sodium violently reacts with water in a type of **displacement reaction**. However, if you were to repeat the same experiment with 100 g of pure gold, nothing would happen at all. Evidently, there is a difference in reactivity between these metals, but why?

To answer this question, we must conduct some important *practical investigations* that compares the reactivity of a range of metals in different conditions. By doing so, we can *rank* the reactivity of metals and attempt to theoretically rationalize our results.

THE MECHANISM OF REACTION: OXIDATION AND REDUCTION

Oxidation – Reduction reactions (or **redox** reactions) are reactions in which there is

……

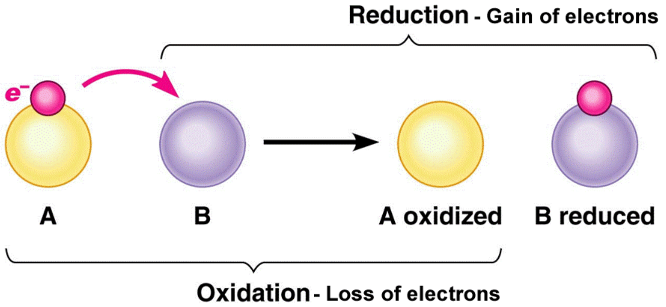
In earlier modules, you have come to understand that when metals react, they form positive ions called cations (i.e. Na^+, Ca^{2+}). To gain this positive charge, these metals must have *lost* their valence electrons in the **S** subshell to achieve a stable electronic configuration. This is the general mechanism of reaction for metals.

These names are simply labels for reaction processes *you have already encountered*.

In reactions involving metals, the metal ………………………………………………… in a process known as

…………………………….

Meanwhile, the other reactant ………………………………………………… these electrons in a process known as …………………………….



This electron transfer can remember by the mnemonic "OIL RIG":

Oxidation

I s

Loss of electrons

Reduction

Is

Gain of electrons

The general equation for the process of oxidation is given below:

$$M_{(s)} \rightarrow M^{+} + {}^{0}_{-1}e$$

Whilst the focus of this lesson is the reactivity of metals, similar equations to the one above will be used to describe chemical processes. It is important that you are able to interpret the meaning of these equations as you will be required to reproduce similar equations in exams when describing redox chemical processes. These concepts will be covered in greater detail in the following two lessons.

REACTIVITY OF METALS IN WATER

Some metals *may react* with water according to the general word equation below:

$$Very\ Reactive\ Metal + Water \rightarrow Metal\ Hydroxide + Hydrogen\ gas$$

As water is not considered a reactive liquid, you would expect only reactive metals to bring about this chemical reaction.

Consider the reaction between sodium metal and water. The neutral species equation is:

$$2Na_{(s)} + 2H_2O_{(l)} \rightarrow 2NaOH_{(aq)} + H_{2(g)}$$

To form aqueous sodium hydroxide, the sodium atom has lost its valence electron and is **oxidized** according this following oxidation **half equation**:

$$Na_{(s)} \rightarrow Na^+_{(aq)} + e^-$$

As the name suggests, a half equation only provides half of the reaction and in the example above, it describes the oxidation process in terms of a chemical equation.

The electron released by sodium is transferred to the hydrogen ions in water which are **reduced** according to this reduction **half equation:**

$$2H_2O_{(l)} + 2e^- \rightarrow H_{2(g)} + 2OH^-_{(aq)}$$

Notice how *two electrons* appear in the reaction above as two sodium atoms react with two water molecules due to the stoichiometry of the balanced equation.

These half equations can be added together to form the **net ionic equation**:

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

In summary, when a metal reacts with water, electrons are transferred from the metal to the water yielding a metal hydroxide and hydrogen gas.

Question 1 (6 marks)

Write the word and balanced chemical equations for the following reactions given that these metals react with water.

a) Potassium and water

……

……

……

……

b) Lithium and water

……

……

……

……

c) Calcium and water

……

……

……

……

Reactivity of Metals With Dilute Acids

The reaction of a metal with an acid can be represented using this general word equation you have encountered in the previous lesson:

$$Metal + Acid \rightarrow Metallic\ Salt + Hydrogen\ gas$$

When a metal reacts with an acid, there is an **electron transfer** involved from the metal analogous to the reaction of a reactive metal with water. As acids are more reactive solutions than water, we would expect metals with *lower reactivity* to also participate in this reaction.

For example, consider the reaction of magnesium with dilute hydrochloric acid.

The **balanced** chemical equation is:

$$Mg_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$$

We can expand this reaction equation to generate a **full ionic equation:**

$$Mg_{(s)} + 2H^+_{(aq)} + 2Cl^-_{(aq)} \rightarrow Mg^{2+}_{(aq)} + H_{2(g)} + 2Cl^-_{(aq)}$$

If we look more closely, there are two separate "half reactions" like before. Magnesium metal has lost its two valence electrons and has been **oxidized** according to the **oxidation half equation:**

$$Mg_{(s)} \rightarrow Mg^{2+}_{(aq)} + 2e^-$$

These two electrons are transferred to *two* hydrogen ions in the process of **reduction** according to the **reduction half equation:**

$$2H^+_{(aq)} + 2e^- \rightarrow H_{2(g)}$$

These half equations can be added together to form the **net ionic equation**:

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

Notice how the chloride ions are missing in the above expression as they are spectator ions that do not participate in the reaction.

In summary, when a metal reacts with an acid, electrons are transferred from the metal to the hydrogen ion of the acid. The metal is oxidized (loss of electrons) to its ion, and the hydrogen ion is reduced (gain of electrons) to hydrogen gas.

Question 2 (7 marks)

Write the balanced equation for the following reactions with correct states:

a) Aluminium with sulfuric acid (2 marks)

....

....

....

....

b) Magnesium with hydrochloric acid (2 marks)

....

....

....

....

c) A student added a small quantity of finely divided magnesium metal into a beaker containing 20 mL of 2M HCl acid and recorded the temperature change during the reaction process. It was determined that the maximum temperature reached was 60˚C.

Explain why the temperature of the beaker increased with reference to the chemical processes occurring and general principles of reactivity. (3 marks)

....

....

....

....

....

....

....

....

....

....

REACTIVITY OF METALS WITH OXYGEN

Metals react with oxygen to form metal oxides. However, the ease at which a particular metal will form its corresponding oxide depends largely on the reaction conditions and the reactivity of the metal itself.

The general word equation is the same for all metal oxides:

$$Metal + Oxygen \rightarrow Metal\ Oxide$$

For example, magnesium metal reacts with oxygen to form magnesium oxide:

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

The metal is **oxidized** as shown in the oxidation half equation to release electrons:

$$Mg_{(s)} \rightarrow Mg^{2+}_{(s)} + 2e^-$$

Meanwhile, the oxygen gas is reduced to oxygen ions as shown in the reduction half equation:

$$O_{2(g)} + 4e^- \rightarrow 2O^{2-}_{(s)}$$

In summary, metals react with oxygen to form metallic oxides.

Question 3 (2 marks)

Write the balanced equation for the following reactions:

a) The formation of a surface oxide layer of chromium (III) oxide

……

……

b) The combustion of zinc to form zinc oxide

……

……

SUMMARY

- The reactions of metals with water, acids and oxygen gas all involve the transfer of electrons *from the* metal to the other reactant.
- The metal is **oxidised** in this process and the other reactant is **reduced**
- If these reactions release heat energy into the surroundings, they are classed as **exothermic reactions.**
- Not all metals have the same level of reactivity.
- The more reactive the metal, the *easier* it is for the metal to give up its valence electrons and oxidise.
- These empirical observations can be used to **order the reactivity** of a range of metals relative to each other.

This is an important page that you will need to refer to in later lessons.

THE REACTIVITY OF A RANGE OF METALS

An important aspect of conducting experimental work is to *compare* the experimental data to secondary sources. Below are three sets of tables that describe the reactivity of a range of metals in the conditions we have studied so far.

	K Na Li Ba Ca	Mg Al Zn Fe	Sn Pb Cu Hg	Ag Au
Reactivity with oxygen	Burns very readily at room temperature	React slowly at room temperature; burns readily if heated in powdered form	At room temperature, surface becomes coated in oxide layer. React slowly when heated	No reaction

	K Na Li Ba Ca	Mg	Al Zn Fe	Sn Pb Cu Hg Ag Au
Reactivity with water	Reacts with cold water	Reacts with hot water	Reacts with steam at red heat	No reaction

	K Na Li	Ba Ca Mg	Al Zn Fe Sn	Pb Cu Hg Ag Pt Au
Reactivity with dilute acid	Extremely rapid bubbling. Hydrogen produced. May ignite	Rapid bubbling with hydrogen produced	Slow bubbling	No reaction

AN ACTIVITY SERIES FOR METALS

Using these simple tests, scientists were able identify which groups of elements were reactive. It is no mystery that group I and II metals such as potassium, sodium, lithium, barium and calcium are very reactive whereas late transition metals such as gold and silver are not. By combining these observations in conjunction with other tests, an activity series for metals could be justified.

Most reactive

K	Potassium
Na	Sodium
Li	Lithium
Ba	Barium
Ca	Calcium
Mg	Magnesium
Al	Aluminium
C	*Carbon*
Zn	Zinc
Fe	Iron
Sn	Tin
Pb	Lead
H	*Hydrogen*
Cu	Copper
Ag	Silver
Pt	Platinum
Au	Gold

Least reactive

This list above outlines the relative reactivity's of common metals. Note that the non-metals hydrogen and carbon are included *for a comparison*. As the metal's reactivity **increases**, its ease of oxidation **increases** making the comparison of reaction rates in acid, oxygen and water a justifiable method of determining metal reactivity.

Question 4 (4 marks)

a) Identify the metal which is the least reactive:

i. Aluminium, Magnesium, Zinc

………………………………………………………………………………………………

ii. Calcium, Copper, Lead

………………………………………………………………………………………………

b) Identify the metal which is the most reactive:

i. Silver, Copper, Gold

………………………………………………………………………………………………

ii. Lithium, Potassium, Aluminium

………………………………………………………………………………………………

Question 5 (2 marks)

Which metals react with cold dilute hydrochloric acid but not cold water? Explain this referring to their relative reactivity.

………………………………………………………………………………………………

………………………………………………………………………………………………

………………………………………………………………………………………………

………………………………………………………………………………………………

Question 6 (2 marks)

Which metals **must not** be stored in containers filled with water, air or acid? Explain why.

………………………………………………………………………………………………

………………………………………………………………………………………………

………………………………………………………………………………………………

………………………………………………………………………………………………

Question 7 (12 marks)

Write balanced equations for the following reactions and describe the environment in which the reaction occurs by using the reactivity series above:

a) Aluminium and oxygen

………

………

b) Tin and water

………

………

c) Potassium with dilute acid

………

………

d) Nickel and steam

………

………

e) Lead and water

………

………

f) Chromium and oxygen

………

………

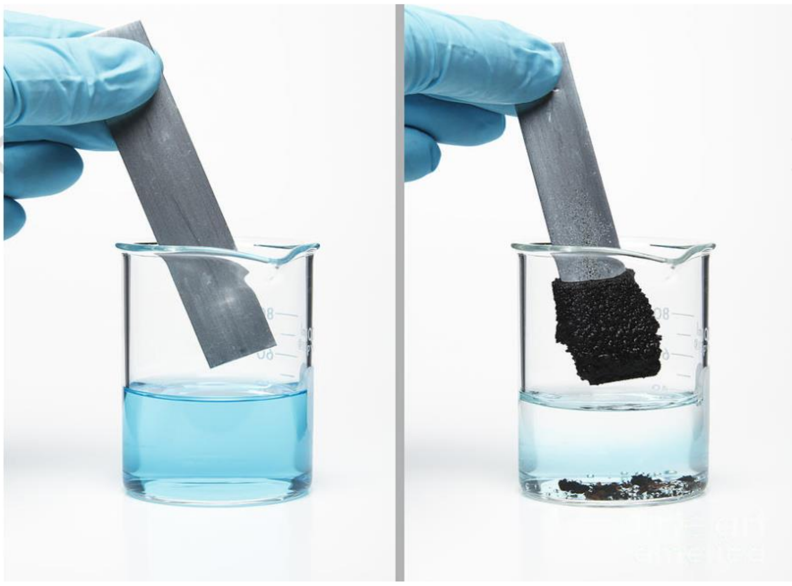
Reactivity of Metals With other Metal Ions in Solution

The order in which metals appear in the activity series is important to predict reactions between metals and other metallic ions in solution. These are known as **metal displacement reactions.**

As we have discovered, a metal's *ease of oxidation* i.e. its ability to lose electrons *increases* with reactivity. It follows that a more reactive metal (in the presence of the ions of a less reactive metal) will be *oxidised* and donate its electrons to the cation of the less reactive metal.

In a displacement reaction: a more ………………………………………………… displaces the ion of a ………………………………………………… metal from solution.

Let's consider the reaction between zinc metal and copper ions. As the activity series indicates, zinc metal is more reactive than copper. If a strip of zinc metal is placed into a solution of copper ions, a spontaneous reaction would follow.



The zinc metal is oxidised and **displaces the copper ions from solution**. As the reaction proceeds, solid copper is deposited on the metal strip as the concentration of zinc ions in solution increases according to the reaction below:

$$Zn_{(s)} + Cu^{2+}_{(aq)} \rightarrow Cu_{(s)} + Zn^{2+}_{(aq)}$$

However, if *reverse the scenario* and place a strip of copper metal in a solution containing zinc ions, no reaction would spontaneously occur. The copper metal **cannot** displace zinc ions from solution as it is a relatively less reactive metal than zinc.

The chemical concept of a displacement reaction can also be used to explain why some metals react with acid while others do not. Iron, magnesium, sodium and calcium are more reactive than hydrogen. Therefore, in the presence of hydrogen ions (acid), these metals will be oxidized and a reaction will occur. However, silver and gold are elements that are *less reactive* than hydrogen, they will not oxidise in an acidic solution.

Question 8 (5 marks)

Determine whether a displacement reaction will occur given the following combination of chemicals. If there is a displacement reaction, write a balanced the chemical equation.

a) $Mg_{(s)} + Ca(NO_3)_{2(aq)} \rightarrow$

……

……

b) $Mg_{(s)} + Mg(NO_3)_{2(aq)} \rightarrow$

……

……

c) $Mg_{(s)} + Zn(NO_3)_{2(aq)} \rightarrow$

……

……

d) $Mg_{(s)} + Cu(NO_3)_{2(aq)} \rightarrow$

……

……

e) $Mg_{(s)} + Ag(NO_3)_{2(aq)} \rightarrow$

……

……

Question 9 (4 marks)

Determine whether a displacement reaction will occur given the following combination of chemicals. If there is a displacement reaction, write a balanced the chemical equation.

a) $Zn_{(s)} + Ca(NO_3)_{2(aq)} \rightarrow$

………

………

b) $Zn_{(s)} + Mg(NO_3)_{2(aq)} \rightarrow$

………

………

c) $Zn_{(s)} + Zn(NO_3)_{2(aq)} \rightarrow$

………

………

d) $Zn_{(s)} + Cu(NO_3)_{2(aq)} \rightarrow$

………

………

Question 10 (3 marks)

Solutions of zinc nitrate and tin(II) nitrate were prepared in a laboratory by a student however without any labels. Name a metal that can be used in to confirm the identity of each solution and explain your choice in terms of the relative reactivity of the metals.

………

………

………

………

………

………