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Content Acknowledgement

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WHY YOU SHOULD STUDY SCIENCE

Science is amazing! It is also one of the toughest subjects at school. Science-y careers are diverse and exciting, but require years of vigorous academic commitment. If it's so hard to get somewhere with Science, why should you study it? Here's our top reasons for getting your nerd on:

1. Be a modern-day hero: The single greatest reason why we should study Science, is to ensure Humanity’s sustainable survival on Earth! Ecosystems are in crisis mode, the planetary weather system is changing rapidly, and humanity is failing to coexist in harmony with other species. World food production has to double in the next thirty years, in order to sustain the growing global population. We are running out of fossil fuels which are critical to the efficiency of our industry, farming and supply chains. Fresh water is becoming increasingly scarce, with many of the World’s greatest rivers no longer running into the sea. Diseases are becoming increasingly resistant to antibiotics. The air in many Indian and Chinese cities are verging on unbreathable. The Great Pacific Garbage Patch has become an unfathomable mass of floating junk that is destroying our oceans. The use of fossil fuels is polluting our air and adding to the Greenhouse Effect.

Before you despair, there is a silver lining: one of these problems can be improved, and even solved, through Science! If you are passionately concerned about this Planet and about a healthy future for Humanity, get stuck into your Science studies and aim for a Science-y career that will equip you to make a difference!

2. Be smart: The study of Science encourages problem-solving tenacity that helps you to understand the world around you. I have always explained to my students that Science illuminates one’s path, and that going through life without Science is similar to driving your car along dark roads - your headlights might light your way forward, but they don’t illuminate the world around you. You travel onwards without ever understanding the context of your journey. Studying Science makes you comfortable with the unknown, and gives you the confidence to say: “I don’t know the answers, but I will find out!”. Science is gracious to naivety but does not condone the apathy of indifference: it allows you to say “I don’t know, but I want to find out”, but does not tolerate the attitude of “I don’t know and I don’t care”.

Science is highly structured, but welcomes change - it constantly adjusts its views based on what is observed. This approach teaches you to evolve your thinking by constantly testing and investigating information, which makes you a well-rounded human being and empowers you with an ethical approach to others: it enables you to discern the difference between your opinions and facts, and to acknowledge the opinions and beliefs of others without immediately accepting or rejecting them.

3. Be adventurous: Science gets you places! I can only speak from my experience - my engineering background, which is firmly rooted in Science, has opened a door to great adventure and exploration. I have worked on four continents and have been exposed to a diversity of incredible experiences that a ‘normal’ office job would never allow. Would you like to work in jungles? Study Natural Sciences. A life of studying volcanoes or auroras, perhaps? Study geosciences. Would you like to ply your mind to solving massive problems and driving innovation? Study engineering! Would you like to work with killer whales? Study zoology!

Science-y careers and research allow you visit places that would not be accessible through other fields of study. Whether you want to go to Antarctica or to outer space, Science is the way to get there.

4. Diversity and flexibility: From dentistry to plasma physics, Science-y careers offer vast opportunities for professional career development and diversification. Engineers are welcomed into the financial sector, due to their problem-solving ability and analytical way of thinking. Many academic physicists teach, perform ground-breaking research and consult private clients in the same work week. Medical professionals diversify into the legal field to become patent attorneys or medical lawyers. However on the flip side, it’s rare for a professional with a ‘non-Science-y’ background to bridge into the Science-based career fields.

5. Inventions: Science-y careers create an intellectual and business environment that is conducive to problem solving and invention. Look at all the exciting inventions of the last twenty years, that have completely transformed our lifestyles. The Internet, the everyday use of GPS, mobile phone technology, PC and touch-screen displays are but a few. This technological progress was made possible due to Science.

Visit online crowdfunding platforms such as Kickstarter and Indiegogo, and appraise the exciting Science-y inventions that are being funded. The tech scene is mushrooming with skunkworks and hackathons that are creating radical innovations. It is an exciting time to be part of Science and technology, and if you want to be at the cusp of making cool things that make a big difference, study Science!

6. Be a modern-day hero (#2): South Africa has a growing deficit of expert Science teachers. If you are passionate about Science, and passionate about making a difference, teaching is a massively rewarding career path that is becoming increasingly lucrative. Remember, supply and demand dictate going rates - if there are fewer expert Science teachers around, the demand for expertise leads to increased fees. Become a Science teacher, a thought leader and a role model!

7. Wealth: More than a fifth of the planet's wealthiest people on the Forbes 2015 list studied an engineering degree, according to a recent survey by the Approved Index platform. A quarter of the Forbes top-hundred have Science as a foundation for their work.

8. Discovery: Science research is a field that allows you to discover the unknown. The deep oceans are unexplored, nanotechnology and photonic crystals have so many secrets, and we're still not sure whether there is any form of life outside near-Earth space. Imagine being the person that publishes a peer-reviewed article to tell the world about a brand new discovery, or a new revelation in our understanding, or a life-altering breakthrough in technology.

This is a call to action for young history-makers, and for a new wave of heroes to save this world and make a difference. I encourage you to become part of it!

James Hayes
Founder – Science Clinic

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<td>Quantitative aspects of chemical change</td>
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<td>80</td>
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<td>Energy and chemical change</td>
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Physics
### TABLE 1: PHYSICAL CONSTANTS

<table>
<thead>
<tr>
<th>Name</th>
<th>Symbol</th>
<th>Value</th>
</tr>
</thead>
<tbody>
<tr>
<td>Acceleration due to gravity</td>
<td>$g$</td>
<td>$9.8 \text{ m/s}^2$</td>
</tr>
<tr>
<td>Gravitational constant</td>
<td>$G$</td>
<td>$6.67 \times 10^{-11} \text{ N\cdot m}^2/\text{kg}^2$</td>
</tr>
<tr>
<td>Radius of Earth</td>
<td>$R_E$</td>
<td>$6.38 \times 10^6 \text{ m}$</td>
</tr>
<tr>
<td>Coulomb's constant</td>
<td>$k$</td>
<td>$9 \times 10^9 \text{ N\cdot m}^2/\text{C}^2$</td>
</tr>
<tr>
<td>Mass of the Earth</td>
<td>$M$</td>
<td>$5.98 \times 10^{24} \text{ kg}$</td>
</tr>
<tr>
<td>Electron mass</td>
<td>$m_e$</td>
<td>$9.11 \times 10^{-31} \text{ kg}$</td>
</tr>
<tr>
<td>Charge on electron</td>
<td>$e$</td>
<td>$1.6 \times 10^{-19} \text{ C}$</td>
</tr>
<tr>
<td>Speed of light in vacuum</td>
<td>$c$</td>
<td>$3.0 \times 10^8 \text{ m/s}$</td>
</tr>
<tr>
<td>Acceleration due to gravity</td>
<td>$a$</td>
<td>$g \times (1\text{ m/s}^2 + 2\text{ ax})$</td>
</tr>
</tbody>
</table>

### Table 2: Formulas

#### MOTION

<table>
<thead>
<tr>
<th>Formula</th>
<th>Description</th>
</tr>
</thead>
<tbody>
<tr>
<td>$\Delta x = v_i \Delta t + \frac{1}{2} a \Delta t^2$</td>
<td>Displacement</td>
</tr>
<tr>
<td>$v_f = v_i + a \Delta t$</td>
<td>Final velocity</td>
</tr>
<tr>
<td>$v_i = \sqrt{v_f^2 - 2 a \Delta x}$</td>
<td>Initial velocity</td>
</tr>
</tbody>
</table>

#### FORCE

<table>
<thead>
<tr>
<th>Formula</th>
<th>Description</th>
</tr>
</thead>
<tbody>
<tr>
<td>$F_{\text{net}} = ma$</td>
<td>Net force</td>
</tr>
<tr>
<td>$F = \frac{Gm_1 m_2}{r^2}$</td>
<td>Gravitational force</td>
</tr>
<tr>
<td>$F = \mu N$</td>
<td>Frictional force</td>
</tr>
</tbody>
</table>

---

Information sheets – Paper 1 (Physics)
**Electric Circuits**

- Electric conductor: A conductor that obeys Ohm's law.
- Ohmic conductor: A conductor that obeys Ohm's law.
- Non-Ohmic conductor: A conductor which does not obey Ohm's Law.
- Internal resistance: The resistance of the cell/battery itself.

**Electromagnetism**

- Refers to the use of 1 kilowatt of electricity for 1 hour.
- Kilowatt hour (kWh): The rate at which electrical energy is converted in an electric circuit and is measured in watts (W).

**Electrostatics**

- Electric field: The field of electrical force exerted from electric charge. The direction of the electric field at a point is the direction that a positive test charge would move if placed at that point.
- Electric field at a point: The electric field at a point is directly proportional to the magnitude of the charge and inversely proportional to the square of the distance from the charge.
- Coulomb's Law: The magnitude of the electrostatic force exerted by one point charge on another point charge is 
  \[ F \propto \frac{Q_1 Q_2}{r^2} \]

**Optics**

- Refractive index: The ratio of the speed of light in air to the speed of light in a specific medium.
- Snell's Law: The angle of refraction is proportional to the angle of incidence.
- Newton's Law of Universal Gravitation: The force of gravity that one mass exerts on another is proportional to the product of their masses and inversely proportional to the square of the distance between them.

**Vectors**

- Resultant vector: The vector that has the same effect as the original vectors acting together.
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RESOLVING INTO COMPONENTS

Diagonal vectors can be broken into components. When vectors are broken into the $x$- and $y$-components, we are determining the horizontal ($x$-axis) and vertical ($y$-axis) effect of the vector.

$F_x = F \cos \theta$

$F_y = F \sin \theta$

COMPONENTS ON A SLOPE

When forces act on objects on a slope, it is useful to resolve vectors into components that are parallel ($\parallel$) or perpendicular ($\perp$) components. The most common force resolved into components on a slope is weight ($F_g$).

$F_{g\parallel} = F_g \sin \theta$

$F_{g\perp} = F_g \cos \theta$

CONSTRUCTING FORCE TRIANGLE

When forces are not co-linear, force triangles can be used to determine resultant forces or the equilibrant. When force triangles are formed, basic geometric rules can be used to determine vectors or resultants.

<table>
<thead>
<tr>
<th>Tail-to-head</th>
<th>Parallelogram</th>
<th>Manipulation</th>
</tr>
</thead>
<tbody>
<tr>
<td>Used for consecutive vectors (vectors that occur in sequence).</td>
<td>Used for vectors that act concurrently on the same object. The resultant is the diagonal of a parallelogram that originates from the tail of the vectors.</td>
<td>The vector arrows can be manipulated to form a force triangle to determine the resultant forces or an equilibrant. When manipulating the vector arrows, the following has to remain the same:</td>
</tr>
</tbody>
</table>
| Eg. A boat travels 90 m east, and then moves 50 m north. | Eg. Two tugboats apply a force of 6 000N and 5 000N at bearings of 60° and 120° respectively on a cargo ship. | • Length of arrow (magnitude)  
• Angle of the arrow (direction)  
• The direction of the arrow head |
| ![Tail-to-head diagram](image1) | ![Parallelogram diagram](image2) | ![Manipulation diagram](image3) |

This principle can also be applied to more than 2 vectors taken in order. The resultant is from the tail of the first vector to the head of the last.

Eg. An object is suspended from a ceiling by 2 cables. Below is a free body diagram as well as a force triangle that can be used to calculate the values of $T_1$ and $T_2$.

Eg. Two tugboats apply a force of 6 000N and 5 000N at bearings of 60° and 120° respectively on a cargo ship.
**2D Vectors - Resultant and Equilibrant**

**EQUILIBRANT:** The force that keeps a system in equilibrium.

The equilibrant is equal in magnitude but opposite in direction to the resultant force.

---

### PYTHAGORAS (90° ONLY)

Pythagoras can only be applied to vector triangles that are right angle triangles.

**FOR FINDING SIDES:**

\[ R^2 = x^2 + y^2 \]

**FOR FINDING ANGLES:**

\[ \sin \theta = \frac{o}{h} \quad \cos \theta = \frac{a}{h} \quad \tan \theta = \frac{o}{a} \]

### EXAMPLE:

A boat travels 90 m due east, and then moves 50 m due north. Determine the displacement of the boat.

\[
\begin{align*}
R^2 &= x^2 + y^2 \\
R &= \sqrt{90^2 + 50^2} \\
R &= 102.96 \text{ m}
\end{align*}
\]

\[
\begin{align*}
\tan \theta &= \frac{2}{2} \\
\theta &= \tan^{-1}\left(\frac{50}{90}\right) \\
\theta &= 29.05^\circ
\end{align*}
\]

Remember that \( \theta \) calculated is relative to the \( x \)-axis,

\[ \therefore \text{bearing} = 90^\circ - 29.05^\circ = 60,95^\circ \]

\[ \therefore \text{Displacement} = 102.96 \text{ m at a bearing of 60,95}^\circ \]

---

### SINE-RULE (NON-90°)

The sine-rule rule can be used to determine sides and angles of vector triangles that are not right angle triangles.

**FOR FINDING SIDES:**

\[
\begin{align*}
\frac{a}{\sin(A)} &= \frac{b}{\sin(B)} &= \frac{c}{\sin(C)}
\end{align*}
\]

### EXAMPLE:

An object is suspended from a ceiling by two cords as shown in the diagram below. Determine the magnitude of \( T_1 \) and \( T_2 \).

\[
\begin{align*}
F_g &= mg \\
3 \text{ kg} &= (3)(9.8) \\
F_g &= 29.4 \text{ N}
\end{align*}
\]

\[
\begin{align*}
\frac{a}{\sin A} &= \frac{b}{\sin B} &= \frac{c}{\sin(C)} \\
\frac{T_1}{\sin 55^\circ} &= \frac{29.4}{\sin 95^\circ} \\
T_1 &= \sin 55^\circ \times \frac{29.4}{\sin 95^\circ} \\
T_1 &= 24.18 \text{ N}
\end{align*}
\]

\[
\begin{align*}
\frac{a}{\sin A} &= \frac{b}{\sin B} &= \frac{c}{\sin(C)} \\
\frac{T_2}{\sin 30^\circ} &= \frac{29.4}{\sin 95^\circ} \\
T_2 &= \sin 30^\circ \times \frac{29.4}{\sin 95^\circ} \\
T_2 &= 14.76 \text{ N}
\end{align*}
\]
**2D VECTORS - RESULTANT AND EQUILIBRANT**

**CONSTRUCTION**

By accurately constructing vector diagrams, the resultant or equilibrant can be determined by measurement. The diagram has to be drawn using a ruler and protractor. The magnitude, represented by the length of the line, is drawn using an appropriate scale, for example:

- 1 cm = 1 N
- 1 mm = 1 N

**EXAMPLE:**

Two forces are applied on an object as shown in the diagram below. Determine the resultant force by construction.

![Diagram](image)

**COMPONENT ADDITION**

The resultant of diagonal forces can be determined using Pythagoras by determining the x-resultant and y-resultant first. This is especially useful for determining resultants when more than 2 forces act on an object and a force triangle cannot be used.

**Steps to determine resultant using component method:**

1. Determine the x- and y-components of each force.
2. Determine the x- and y-resultants of components.
3. Determine the resultant using Pythagoras.
4. Determine the angle using trigonometric principles.

**EXAMPLE:**

Three forces act on an object as shown in the diagram below. Determine the resultant force on the object.

1. Determine the x- and y-components of each force.
   - **11N force:**
     
     \[
     F_x = F \cos \theta = 11 \cos 70 = 3.76 \text{ N right (90°)}
     \]
     \[
     F_y = F \sin \theta = 11 \sin 70 = 10.34 \text{ N up (0°)}
     \]
   - **30N force:**
     
     \[
     F_x = F \cos \theta = 30 \cos 40 = 22.98 \text{ N left (270°)}
     \]
     \[
     F_y = F \sin \theta = 30 \sin 40 = 19.28 \text{ N down (180°)}
     \]
   - **20N force:**
     
     \[
     F_x = F \cos \theta = 20 \cos 35 = 16.38 \text{ N right (90°)}
     \]
     \[
     F_y = F \sin \theta = 20 \sin 35 = 11.47 \text{ N down (180°)}
     \]

2. Determine the x- and y-resultants of components.
   - **Take left (270°) as positive**
     
     \[
     F_x = -3.76 + 22.98 - 16.38
     \]
     \[
     F_y = -10.34 + 19.28 + 11.47
     \]
     \[
     F_x = 2.84 \text{ N left (270°)}
     \]
     \[
     F_y = 20.41 \text{ N down (180°)}
     \]

3. Find resultant - Pythagoras.
   \[
   R^2 = x^2 + y^2
   \]
   \[
   R = \sqrt{2.84^2 + 20.41^2}
   \]
   \[
   \theta = \tan^{-1} \frac{20.41}{2.84}
   \]
   \[
   R = 20.61 \text{ N}
   \]
   \[
   \theta = 82.08°
   \]

**RESULTANT:**

- Resultant = 20.61 N at a bearing of 187.92°

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Newton’s Laws of Motion

Forces

<table>
<thead>
<tr>
<th>Non-contact force</th>
<th>Contact force</th>
</tr>
</thead>
<tbody>
<tr>
<td>A force exerted between objects over a distance without physical contact.</td>
<td>A force exerted between objects that are in contact with each other.</td>
</tr>
</tbody>
</table>

- Electrostatic force
- Gravitational force ($F_g$)
- Magnetic force

A force is a push or a pull action exerted on an object. This action can be exerted while objects are in contact (contact force) or over a distance (non-contact force).

Because forces have magnitude and direction, they are vectors. Force is measured in newton (N). 1 N is the force required to accelerate a 1 kg object at 1 m·s$^{-2}$ in the direction of the force. We can therefore say that $1 \text{ N} = 1 \text{ kg} \cdot \text{m} \cdot \text{s}^{-2}$.

Non-contact force:
A force exerted between objects over a distance without physical contact.

Contact force:
A force exerted between objects that are in contact with each other.

Electrostatic force
- Applied force ($F_A$)

Gravitational force ($F_g$)
- Tension ($T$ or $F_T$)

Magnetic force
- Friction ($F_f$ or $f_s/f_k$)
- Normal force ($F_N$)

Normal force ($F_N$)

The perpendicular force exerted on by a surface on an object in contact with it.

The normal force is equal to the perpendicular component of gravity if there are no other forces acting on the object.

If alternative forces act on the object, the normal force will change depending on the direction and magnitude of the applied force.

Objects suspended from a rope/string/cable have no normal force, as there is no surface on which the object rests. The tension is equal to the perpendicular component of gravity if there are no other forces acting on the object.

Fric%on ($F_f$) or $f_s/f_k$
Frictional force due to a surface is the force that opposes the motion of an object in contact with it.

Friction is the parallel component of the contact force on an object by the surface on which it rests. The friction between the contact surfaces are determined by the properties of that surface. The coefficient of friction ($\mu_s/\mu_k$) is a description of the roughness of the surface. The rougher the surface, the greater the coefficient of friction.

Static friction ($f_s$)
Static friction is the frictional force on a stationary object that opposes the tendency of motion of the object. The magnitude of the static friction will increase as the parallel component of the applied force is increased, until maximum static friction is reached. $f_{s\text{max}}$ is the magnitude of friction when the object just starts to move.

Kinetic friction ($f_k$)
Kinetic friction is the frictional force on a moving object that opposes the motion of the object. The magnitude of the kinetic friction is constant for the specific system at all velocities greater than zero, and irrespective of the applied force.

$F_N + F_A\sin \theta = F_g$
$F_N - F_g - F_A\sin \theta = 0$
$F_N = F_g + F_A\sin \theta$

If the applied force is greater than the maximum static friction, the object will start to move.

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Newton’s First Law of Motion

An object continues in a state of rest or uniform (moving with constant) velocity unless it is acted upon by an unbalanced (net or resultant) force.

Newton’s First Law is due to inertia: the resistance of an object to change its state of rest of motion.

$F_{\text{net}} = 0 \text{ N} \quad \text{and} \quad a = 0 \text{ m} \cdot \text{s}^{-2}$

A 3 kg object moves up an incline surface at an angle of 15° with a constant velocity. The coefficient of friction is 0.35. Determine the magnitude of the applied force.

Importance of wearing safety belts:

According to Newton’s First Law, an object will remain in motion at a constant velocity unless a non-zero resultant force acts upon it. When a car is in an accident and comes to a sudden stop, the person inside the car will continue with a constant forward velocity. Without a safety belt, the person will make contact with the windscreen of the car, causing severe head trauma. The safety belt acts as an applied force, preventing the forward motion of the person.

Newton’s Second Law of Motion

When a net force is applied to an object of mass, it accelerates in the direction of the net force. The acceleration is directly proportional to the net force and inversely proportional to the mass.

Newton's Second Law is dependent on the resultant force acting on the object.

$F_{\text{net}} = ma \quad \text{and} \quad a \neq 0 \text{ m} \cdot \text{s}^{-2}$

A 20 N force is applied to a 5 kg object. The object accelerates up a frictionless incline surface at an angle of 15°. Determine the acceleration of the object.

Effect of Newton’s Second Law on overloading:

According to Newton’s Second Law, the acceleration of an object is directly proportional to the applied force and inversely proportional to the mass of the object. If a vehicle is overloaded, the stopping distance will increase which can lead to serious accidents. When brakes are applied, the force (friction) remains the same, but the increase in mass causes a decrease in negative acceleration, increasing the time (and distance) it takes for the vehicle to stop.

Force pairs properties:
- Equal in magnitude
- Opposite in direction
- Acts on different objects (and therefore DO NOT CANCEL each other out)

NOTE: The force pairs shown here are gravitational forces.
Gravity and Normal force are NOT force pairs.
**Horizontal**

The vertical resultant = 0 N.
The horizontal resultant determines acceleration.

**Pulled at an angle**

Horizontal:

\[ F_{\text{net}} = ma \]

\[ F_{Ax} + (-F_f) = ma \]

\[ (-F_g) + F_N + F_{Ay} = 0 \]

**Pushed at an angle**

Horizontal:

\[ F_{\text{net}} = ma \]

\[ F_{Ax} + (-F_f) = ma \]

\[ (-F_g) + F_N + F_{Ay} = 0 \]

**Vertical**

\[ F_{\text{net}} = 0 \]

\[ F_f + (-F_N) + F_{Ay} = 0 \]

**Suspended**

Horizontal resultant = 0 N.
Vertical resultant determines acceleration.

**Lift stationary/constant velocity**

Vertical:

\[ F_{\text{net}} = 0 \]

\[ F_g + (-F_T) = 0 \]

**Lift accelerating**

Vertical:

\[ F_{\text{net}} = ma \]

\[ F_g + (-F_T) = ma \]

**Lift in freefall (cable snap)**

Vertical:

\[ F_{\text{net}} = ma \]

\[ F_g = ma \]

Acceleration will be in the direction of the greatest force.

**Slopes**

The perpendicular (⊥) resultant = 0 N.
The parallel (∥) resultant determines acceleration.

**All examples:**

\[ F_{g\parallel} = F_g \sin \theta \]

\[ F_{g\perp} = F_g \cos \theta \]

**Remember:** Use components of weight.

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**Connected objects**
Do separate free body diagrams for each object.
The velocity and acceleration of all objects are equal in magnitude and direction.
Applied forces are applied to only one object at a time.
Simultaneous equations for acceleration and tension are sometimes needed.

**REMEMBER:**
- Ropes/cables: The tension forces on the objects are the same in magnitude but opposite in direction.
- Touching objects: Newton’s Third Law

**Same axis**
Can be horizontal (multiple objects on a surface) or vertical (multiple suspended objects).
The velocity and acceleration of all objects are equal in magnitude and direction.

**Multiple axes**
Horizontal (objects on a surface) **AND** vertical (suspended objects).
The velocity and acceleration of all objects are equal in magnitude **NOT** DIRECTION.

**Vector direction on multiple axes**
Clockwise:
- Right and Down positive
- Left and Up negative

**OR**
Anti-clockwise:
- Left and Up positive
- Right and Down negative

---

**Newton’s Laws of Motion**

**Objects attached by rope/cable**

**Horizontal:**
\[ \sum F_x = 0 \]
\[ F_{net} = ma \]
\[ (-F_f) + F_T = ma \]

**Objects in contact**

**Horizontal:**
\[ \sum F_x = 0 \]
\[ F_{net} = ma \]
\[ F_A(-F_x) + (-F_f) = ma \]

**Multiple axes**

**In these examples,**
clockwise is positive:
- Right positive
- Down positive

**Horizontal:**
\[ \sum F_x = 0 \]
\[ F_{net} = ma \]
\[ F_{T1}(-F_f) + (-F_f) = ma \]

**Vertical:**
\[ \sum F_y = 0 \]
\[ F_{net} = ma \]
\[ F_{T1}(-F_f) + (-F_f) = ma \]

---

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Every particle in the universe attracts every other particle with a force which is directly proportional to the product of their masses and inversely proportional to the square of the distance between their centres.

\[ F = \frac{Gm_1m_2}{r^2} \]

- **F** = force of attraction between objects (N)
- **G** = universal gravitational constant \( \left( 6.67 \times 10^{-11} \text{ N}\cdot\text{m}^2\text{kg}^{-2} \right) \)
- **m** = object mass (kg)
- **r** = distance between object centers (m)

A uniform sphere of matter attracts a body that is outside the shell as if all the sphere’s mass was concentrated at its center.

Thus, the distance is determined between the centers of the two bodies.

**NOTE:**
- The radius of object (man) on the earth is negligibly small.
- The radius of object (moon) on the earth is

<table>
<thead>
<tr>
<th>RATIOS</th>
</tr>
</thead>
<tbody>
<tr>
<td>1. Write out the original formula.</td>
</tr>
<tr>
<td>2. Manipulate unknown as subject.</td>
</tr>
<tr>
<td>3. Substitute changes into formula (Keep symbols!).</td>
</tr>
<tr>
<td>4. Simplify ratio number.</td>
</tr>
<tr>
<td>5. Replace original formula with unknown symbol.</td>
</tr>
</tbody>
</table>

**EXAMPLE:**

Two objects, \( m_1 \) and \( m_2 \), are a distance \( r \) apart and experience a force \( F \). How would this force be affected if:

a) One mass is doubled and the distance between the masses is halved?

\[ F = \frac{Gm_1m_2}{r^2} \]  
Write out the formula

\[ F = \frac{G(2m_1)m_2}{(\frac{r}{2})^2} \]  
Substitute changes into formula

\[ F = \frac{\frac{1}{2} Gm_1m_2}{\frac{r^2}{4}} \]  
Simplify ratio number

\[ F = 4 \left( \frac{Gm_1m_2}{r^2} \right) \]  
Replace original formula

\[ \therefore F_{new} = 8 F \]

b) Both the two masses as well as the distance are doubled?

\[ F = \frac{Gm_1m_2}{r^2} \]  
Write out the formula

\[ F = \frac{G(2m_1)(2m_2)}{(2r)^2} \]  
Substitute changes into formula

\[ F = \frac{4 Gm_1m_2}{4 r^2} \]  
Simplify ratio number

\[ F = 1 \left( \frac{Gm_1m_2}{r^2} \right) \]  
Replace original formula

\[ \therefore F_{new} = 1 F \]

**DETERMINING GRAavitational ACCELERATION (g)**

\[ F = m_{object}g \]  
\[ F = \frac{Gm_{object}m_{Planet}}{r_{Planet}^2} \]

\[ m_{object} = \frac{Gm_{Planet}r_{Planet}^2}{F} \]

\[ g = \frac{Gm_{Planet}}{r_{Planet}^2} \]

\[ \therefore g = \frac{F}{m_{object}} \]

Therefore the gravitational acceleration of an object only depends on the mass and radius of the planet. Object mass is irrelevant!

**CALCULATIONS**

The force can be calculated using \( F = \frac{Gm_1m_2}{r^2} \)

**EXAMPLE:**

The earth with a radius of \( 6,38 \times 10^3 \) km is 149,6 \( \times 10^8 \) km away from the sun with a radius of 696 342 km. If the earth has a mass of \( 5,97 \times 10^{24} \) kg and the sun has a mass of \( 1,99 \times 10^{30} \) kg, determine the force between the two bodies.

\[ F = \frac{Gm_1m_2}{r^2} \]

\[ F = \frac{6,67 \times 10^{-11} \times 5,97 \times 10^{24} \times 1,99 \times 10^{30}}{(1,5 \times 10^{11})^2} \]

\[ F = 3,52 \times 10^{22} \text{ N attraction} \]

The force of gravitational attraction is a vector, therefore all vector rules can be applied:

- Direction specific
- Can be added or subtracted

**Take right as positive:**

\[ F_{net \ on \ satellite} = F_{\text{on satellite}} + F_{\text{on satellite}} \]

\[ = \left( \frac{Gm_{earth}m_s}{r_{satellite}^2} \right) + \left( \frac{Gm_{sun}m_s}{r_{sun}^2} \right) \]

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**ELECTROSTATICS**

**PRINCIPLE OF CONSERVATION OF CHARGE**
- \( Q_{\text{new}} = \frac{Q_1 + Q_2}{2} \)

- The principle of conservation of charge states that the total charge is conserved during an electrical process.

**PRESSURE**
- \( P = \frac{F}{A} \)

**Coulomb’s Law**
- The force between two charges is directly proportional to the product of the charges and inversely proportional to the square of the distance between the charges.

\[
F = \frac{kQ_1Q_2}{r^2}
\]

- \( r = \text{distance between objects (m)} \)
- \( Q = \text{object charge (C)} \)
- \( k = \text{Coulomb’s constant (9 \times 10^9 \text{N} \cdot \text{m}^2/\text{C}^2)} \)
- \( F = \text{force of attraction between objects (N)} \)

**Universal Gravitation**
- In ratio questions, the same process is used as with Newton’s Law of Universal Gravitation.

**EXAMPLE:**
- Two charges experience a force \( F \) when held a distance \( r \) apart. How would this force be affected if one charge is doubled, the other charge is tripled and the distance is halved.

\[
F = \frac{kQ_1Q_2}{r^2} \quad \Rightarrow \quad \text{new} = \frac{k(2Q_1)(3Q_2)}{(\frac{r}{2})^2} = 24F
\]

**CALCULATIONS - Electrostatic force**
- The force can be calculated using \( F = \frac{kQ_1Q_2}{r^2} \), where:
  - \( \theta = \tan^{-1} \frac{F_{AB}}{F_{CB}} \)
  - \( \theta = \tan^{-1} \frac{4.500}{2.800} \)
  - \( \theta = 58.11^\circ \)

\[
F = \frac{kQ_1Q_2}{r^2}
\]

**FLASHBACK!!**
- Prefix Conversion
  - centi– (cC) \( \times 10^{-2} \)
  - milli– (mC) \( \times 10^{-3} \)
  - micro– (µC) \( \times 10^{-6} \)
  - nano– (nC) \( \times 10^{-9} \)
  - pico– (pC) \( \times 10^{-12} \)

**COULOMB’S LAW**
- The force between two charges is directly proportional to the product of the charges and inversely proportional to the distance between the charges squared.

**RATIOS**

In ratio questions, the same process is used as with Newton’s Law of Universal Gravitation.

- Substitute charge magnitude only.
- Direction determined by charge (like repel, unlike attract).
- Both objects experience the same force (Newton’s Third Law of Motion).

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  - \( \theta = 58.11^\circ \)

\[
F = \frac{kQ_1Q_2}{r^2}
\]
**ELECTRIC FIELDS**

An electric field is a region of space in which an electric charge experiences a force. The direction of the electric field at a point is the direction that a positive test charge (+1C) would move if placed at that point.

**Single point charges**

- **Unlike charges**
  - [Diagram showing electric field lines between two unlike charges]
  - Electric field lines point from positive to negative charge.

- **Like charges**
  - [Diagram showing electric field lines between two like charges]
  - Electric field lines point from negative to negative charge.

**ELECTROSTATICS**

Electric field strength at any point in space is the force per unit charge experienced by a positive test charge at that point.

\[
E = \frac{F}{q}
\]

- **Electric field strength** (N·C\(^{-1}\))
- **F** = force (N)
- **q** = charge (C)

**ELECTRIC FIELD STRENGTH**

\[
E = \frac{kQ}{r^2}
\]

- **Electric field strength** (N·C\(^{-1}\))
- **k** = Coulomb's constant (9 \(\times\) 10\(^9\) N·m\(^2\)·C\(^{-2}\))
- **Q** = object charge (C)
- **r** = distance between objects (m)

**EXAMPLE:**

Charge B experiences a force of 2 N due to charge A. Determine the electric field strength at point B.

\[
E = \frac{F}{q} = \frac{2}{5 \times 10^{-6}} = 4 \times 10^5 \text{ N} \cdot \text{C}^{-1}
\]

**DIRECTION:** Direction that point in space (X) would move IF it was positive.

**NOTE:**

Electric field strength is a VECTOR. All vector rules and calculations apply.

(Linear addition, 2D arrangement, resultant vectors, etc.)

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When waves reach a change in medium they can undergo:

**REFLECTION**
Propagation of light from air into glass and back into air.
Law of reflection: The angle of reflection is equal to the angle of incidence. The incident ray, reflected ray and the normal all lie on the same plane.

**DIFFRACTION**
The ability of a wave to spread out in wavefronts as they pass through a small aperture or around a sharp edge.

**REFRACTION**
When light moves between media of different optical densities, the light undergoes refraction.

*Refractive index:* The bending of light as it moves between two media with different optical densities as a result of the change in wave speed while the frequency remains constant.

*Normal:* An imaginary line perpendicular to the surface of an object.

*Angle of Incidence:* Angle between an incident ray and the normal on the surface.

*Angle of Refraction:* Angle between a refracted ray and the normal on the surface.

### Geometric Optics

#### Optical Density and Refractive Index

**Optical density**

Light can travel through a vacuum at $3 \times 10^8 \text{ m/s}$, but slows down as it moves through a transparent medium. The optical density is a measure of the refracting power of a medium. The higher the optical density, the more the light will be refracted or slowed down as it moves through the medium.

The velocity of light through a medium is inversely proportional to the optical density of the medium. The speed of light in the same medium is constant.

The optical density can be quantified as a refractive index. The refractive index is the ratio of the velocity of light in vacuum to its velocity in a specified medium and can be determined by:

$$n = \frac{c}{v}$$

where
- $n$ = refractive index (no unit)
- $c = \text{speed of light in vacuum} = 3 \times 10^8 \text{ m/s}$
- $v = \text{speed of light through medium} \ (\text{m/s})$

**Refractive index**
The refractive index is defined as the ratio of the speed of light in vacuum ($c$) to the speed of light in a material ($v$). The refractive index is directly proportional to the optical density and inversely proportional to the velocity of light through the medium.

<table>
<thead>
<tr>
<th>Medium</th>
<th>Refractive index</th>
</tr>
</thead>
<tbody>
<tr>
<td>Vacuum</td>
<td>1</td>
</tr>
<tr>
<td>Air</td>
<td>1.003 $\approx$ 1</td>
</tr>
<tr>
<td>Water</td>
<td>1.33</td>
</tr>
<tr>
<td>Crown glass</td>
<td>1.50 - 1.62</td>
</tr>
<tr>
<td>Diamond</td>
<td>2.42</td>
</tr>
</tbody>
</table>

The refractive index ($n$) is also a quantitative description of the amount of refraction (bending) that occurs when light moves between media of different refractive indices. The angle of refraction can be determined using Snell’s Law. There is a relationship between the angles of incidence and refraction when light travels through the boundary between two different optical media.

### Snell’s Law

Snell’s Law: The ratio of the angle of incidence and the angle of refraction is equivalent to the reciprocal of the ratio of refractive indices.

$$\frac{n_i}{n_r} = \frac{n_r}{n_i}$$

where
- $n_i$ = refractive index (origin medium)
- $n_r$ = refractive index (refracted medium)
- $\theta_i$ = angle of incidence (origin medium)
- $\theta_r$ = angle of refraction (refracted medium)

**Change in optical density**

<table>
<thead>
<tr>
<th>Change in optical density</th>
<th>Direction of bend</th>
</tr>
</thead>
<tbody>
<tr>
<td>Less dense to more dense</td>
<td>Towards normal</td>
</tr>
<tr>
<td>More dense to less dense</td>
<td>Away from normal</td>
</tr>
</tbody>
</table>

**Example:**

*Light moves from air into glass ($n = 1.55$) at an angle of $25^\circ$ to the normal. Determine the angle of refraction in the glass.*

$$\sin \theta_i = \frac{n_i}{n_r}$$

$$\sin 25^\circ = \frac{1}{1.55}$$

$$\begin{align*}
\theta_r &= 15.82^\circ
\end{align*}$$
CRITICAL ANGLE

Principle

Critical angle: The angle of incidence in the optically denser medium for which the angle of refraction in the optically less dense medium is 90°. The refracted ray travels parallel to the boundary between the media.

\[ n_1 \sin \theta_1 = n_2 \sin 90° \]

The relationship between the incident angle and the critical angle determines what will happen to light when it moves between media from a high optical density to a low optical density.

- **Refraction**
  \( \theta_1 < \theta_{\text{crit}} \)
  The light ray is refracted at a variable angle according to Snell's Law.

- **Refract 90°**
  \( \theta_1 = \theta_{\text{crit}} \)
  The light ray travels along the border between the two media, at an angle of 90° to the normal.

- **Total Internal Reflection (TIR)**
  \( \theta_1 > \theta_{\text{crit}} \)
  The light ray is reflected internally into the incident medium. The angle of reflection is equal to the angle of incidence.

HUYGENS PRINCIPLE

Huygens’ Principle states that every point on a wavefront can serve as a point source for secondary, circular wavelets with a speed equal to that of the propagation of the waves. After a time, t, the new wavefront will be that of a surface tangent to the secondary waves. A straight wavefront is a result of the superposition of infinite circular waves. When a wave passes through a small aperture or past a sharp edge, it curves around the corners of the obstacle or aperture. This phenomena is known as diffraction.

DEGREE OF DIFFRACTION

The degree of diffraction is directly proportional to the wavelength and inversely proportional to the width of the slit according to:

\[ \text{Diffraction} \propto \frac{\lambda}{w} \]

A narrower slit and longer wave produces more of a curved pattern, in other words, more diffraction takes place.

Wavelength | Slit width
---|---
Longer wavelengths are diffracted more, therefore red light will diffract more than violet light.

DIFFRACTION OF LIGHT

When monochromatic light is passed through a narrow slit, a diffraction pattern forms that consists of alternating light and dark bands. The light bands are as a result of constructive interference. The dark bands form as a result of destructive interference.

USING DOMINANT REVISION

Optical fibers are strands of flexible material strands that allow light to be transmitted over long distances using Total Internal Reflection. Optical fibers consist of an inner glass core with a high refractive index, surrounded by a cladding. The cladding has to have a lower refractive index than the core. Multiple optical fibres can be combined to form a fibre optic cable. Fibre optic cables are primarily used in telecommunications and medicine.

Telecommunications
- Light travels faster than current, therefore signal is faster in fibre optic cables than in copper wires.
- Fibre optic cables are less vulnerable to interception and interference.
- Fibre optic cables have very little resale value, making it less likely to be stolen.

Medicine
- Fibre optics are used in endoscopes.
- Endoscopes consist of 2 flexible fibre optic cables— one to produce light and another to return the image.
- Enables doctors to examine a patient through natural openings in the body such as the mouth, limiting the need for surgery.
- Key-hole surgery can be performed, limiting scarring and recovery time.
**INDUCTION OF A MAGNETIC FIELD**

When current passes through a conductor, a magnetic field is induced around the wire.

The direction of the magnetic field can be determined by the right hand thumb rule. For a straight, single wire, point the thumb of your right hand in the direction of the conventional current and your curled fingers will point in the direction of the magnetic field around the wire.

For a wire loop, the magnetic field is the sum of the individual magnetic fields around the single wires at each end. Use the right hand rule for a single wire at each end of the loop.

For a solenoid, curl your fingers around the solenoid in the direction of the conventional current and your thumb will point in the direction of the North pole. This is known as the right hand solenoid rule.

---

**ELECTROMAGNETISM**

**INDUCTION OF AN ELECTRIC CURRENT**

When a magnet is brought close to a metal wire, it causes movement of charge in the wire. As a result, an EMF is induced in the wire. Only a change in magnetic flux will induce a current.

Faraday's law states that the emf induced is directly proportional to the rate of change of magnetic flux (flux linkage).

\[ \varepsilon = \frac{-N \Delta \phi}{\Delta t} \]

\( \varepsilon \) = emf (V)  
\( N \) = number of turns/windings in coil  
\( \Delta \phi \) = change in magnetic flux (Wb)  
\( \Delta t \) = change in time (s)

Faraday’s law is based on the observation that a changing magnetic flux induces an EMF. The EMF is directly proportional to the rate of change of magnetic flux. The magnetic flux is the result of the product of the perpendicular component of the magnetic field and cross-sectional area the field lines pass through.

\[ \phi = BA \cos \theta \]

\( \phi \) = magnetic flux (Wb)  
\( B \) = magnetic flux density (T)  
\( A \) = area (m²)  
\( \theta \) = angle between magnetic field line and normal

---

**INCREASING THE INDUCED EMF**

- Increase the rate of change of magnetic flux, i.e. decrease the time it takes to change the flux, i.e. increase speed of movement.
- Increase the number of loops in the coil.
- Increase the strength of the magnet.
- Increase the surface area of the loops in the coil.
- Increase the change of flux by changing the angle, \( \theta \), from a minimum of 0° to a maximum of 90°.

---

**ENVIRONMENTAL IMPACT OF OVERHEAD CABLES**

- Birds fly into power lines as they cannot see them from a distance.
- This places birds at risk of becoming extinct because of the increase in unnatural mortality.
- Trees fall onto power lines and can cause fires to erupt.
- Trees have to be cut down to make space for the power lines.
- There is no evidence that the electromagnetic effect of the power lines has any negative impact on people and the surroundings as the strength of the field is low.
- The electromagnetic effect of the power line can disrupt radio signals and for emergency services, this can be a major problem.

---

**DIRECTION OF INDUCED CURRENT**

As a bar magnet moves into a solenoid the needle of the galvanometer is deflected away from the 0 mark. As the bar magnet is removed, the needle deflects in the opposite direction. The magnetic energy is converted to electrical energy. The direction of the induced current can be determined using Lenz’s law.

Lenz’s Law states that the induced current flows in a direction so as to set up a magnetic field to oppose the change in magnetic flux.
**Electricity**

**SERIES**
- If resistors are added in series, the total resistance will increase and the total current will decrease provided the emf remains constant.

**CURRENT** is the rate of flow of charge.

\[ I = \frac{Q}{t} \]

\( I \) is the current strength, \( Q \) the charge in coulombs and \( t \) the time in seconds. SI unit is ampere (A).

**RESISTANCE** is the material’s opposition to the flow of electric current.

\[ V = IR \]

\( V \) is the electrical resistance of the conducting material, resisting the flow of charge through it.

Resistance (R) is the quotient of the potential difference (V) across a conductor and the current (I) in it. The unit of resistance is called the ohm (Ω).

**POTENTIAL DIFFERENCE** (p.d.) is the work done per unit positive charge to move the charge from one point to another. It is often referred to as voltage.

\[ V = \frac{W}{Q} \]

\( V \) is Potential difference in V (volts), \( W \) is Work done or energy transferred in J (joules) and \( Q \) is Charge in C (coulombs).

---

**PARALLEL**
- If resistors are added in parallel, the total resistance will decrease and the total current will increase, provided the emf remains constant.

**CALCULATIONS (NO INTERNAL RESISTANCE)**

**Series circuit**

\[ \frac{1}{R_{eq}} = \frac{1}{R_1} + \frac{1}{R_2} + \frac{1}{R_3} \]

\( R_{eq} \) is the effective resistance of the circuit.

\[ R_{eq} = \frac{1}{\frac{1}{R_1} + \frac{1}{R_2} + \frac{1}{R_3}} \]

**Parallel circuit**

\[ \frac{1}{R_{eq}} = \frac{1}{R_1} + \frac{1}{R_2} \]

\[ R_{eq} = \frac{1}{\frac{1}{R_1} + \frac{1}{R_2}} \]

\( R_{eq} \) is the total resistance.

\[ R_{eq} = \frac{R_1 R_2}{R_1 + R_2} \]

\( R_{eq} \) is the total resistance when resistors are added in series.

\[ R_{eq} = \frac{1}{\frac{1}{R_1} + \frac{1}{R_2}} \]

**Combination circuits**

Consider the circuit given. (Internal resistance is negligible)

Calculate:

- a) the effective resistance of the circuit.
- b) the reading on ammeter A1.
- c) the reading on voltmeter V1.
- d) the reading on ammeter A2.

**Calculations**

\[ R_{eq} = \frac{R_1 R_2}{R_1 + R_2} \]

\[ R_{eq} = \frac{1}{\frac{1}{R_1} + \frac{1}{R_2}} \]

\[ R_{eq} = \frac{1}{\frac{1}{R_1} + \frac{1}{R_2}} \]

\( R_{eq} \) is the total resistance when resistors are added in parallel.

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**OHM’S LAW**

Current through a conductor is directly proportional to the potential difference across the conductor at constant temperature.

\[ R = \frac{V}{I} \]

Proof of Ohm’s law:

The current in the circuit is changed using the rheostat, thus current is the independent variable and potential difference is the dependent variable. It is important that the temperature of the resistor is kept constant.

The resultant graph of potential difference vs current indicates if the conductor is ohmic or non-ohmic.

<table>
<thead>
<tr>
<th>Ohmic conductor</th>
<th>Non-ohmic conductor</th>
</tr>
</thead>
<tbody>
<tr>
<td>An Ohmic conductor is a conductor that obeys Ohm’s law at all temperatures.</td>
<td>A Non-ohmic conductor is a conductor that does not obey Ohm’s law at all temperatures.</td>
</tr>
<tr>
<td>Constant ratio for ( \frac{V}{I} ).</td>
<td>Ratio for ( \frac{V}{I} ) change with change in temperature.</td>
</tr>
<tr>
<td>E.g. Nichrome wire</td>
<td>E.g. Light bulb</td>
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</tbody>
</table>

**ELECTRICITY**

**INTERNAL RESISTANCE**

The potential difference across a battery not connected in a circuit is the emf of the battery.

Emf is the total energy supplied per coulomb of charge by the cell.

When connected in a circuit the potential difference drops due to the internal resistance of the cells.

In reality all cells have internal resistance \( r \). Internal resistors are always considered to be connected in series with the rest of the circuit.

**Determining the emf and internal resistance of a cell**

**POWER**

Power is the rate at which work is done.

\[ P = \frac{W}{\Delta t} \]
\[ P = I^2R \]
\[ P = VI \]
\[ P = \frac{V^2}{R} \]
\[ P = \text{power (W)} \]
\[ W = \text{work (J)} \]
\[ \Delta t = \text{time (s)} \]
\[ I = \text{current (A)} \]
\[ V = \text{potential difference (V)} \]
\[ R = \text{resistance (Ω)} \]

**Example**

A geyser produces 1200 W of power. Calculate the cost of having the geyser switched on for 24 hours, if the price of electricity is 85 c per unit.

\[ \text{Cost} = \text{power \times time \times cost per unit} \]
\[ = (1200 \times 24) \times 85 \]
\[ = 24,480 \]

Cost of electricity = power \times time \times cost per unit

**COST OF ELECTRICITY**

Electricity is paid for in terms of the amount of energy used by the consumer.

Cost of electricity = power \times \text{time} \times \text{cost per unit}

\[ \text{Note: } P \text{ in kW, } t \text{ in hours} \]

**CALCULATIONS (WITH INTERNAL RESISTANCE)**

\[ \epsilon = V_{\text{ext}} + V_{\text{int}} \]

\[ V_{\text{int}} = IR \]

\[ V_{\text{ext}} = IR_{\text{ext}} \]

\[ \epsilon = |R_{\text{ext}} + r| \]

**EXAMPLE:**

When the switch is open the voltmeter reading is 12 V.

When the switch is closed, the current through the 3 Ω resistor is 1 A.

\[ R_{\text{bo}} = R_{\text{3Ω}} + R_{\text{4Ω}} \]
\[ = 5 + 3 \]
\[ = 8 \text{ Ω} \]

\[ R_{\text{bo}} = R_{\text{3Ω}} \]
\[ = 5 \text{ Ω} \]
\[ = R_{\text{3Ω}} \]
\[ = \frac{1}{2} \]

\[ \frac{1}{R_{\text{bo}}} = \frac{1}{R_{\text{3Ω}}} = \frac{1}{2} \]

\[ R_{\text{bo}} = R_{\text{4Ω}} \]
\[ = 4 \text{ Ω} \]
\[ = \frac{1}{2} \]

\[ R_{\text{bo}} = R_{\text{4Ω}} \]
\[ = 4 \text{ Ω} \]
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\[ R_{\text{bo}} = R_{\text{4Ω}} \]
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\[ = \frac{1}{2} \]
Chemistry
TABLE 1: PHYSICAL CONSTANTS

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<th>NAME/NAAM</th>
<th>SYMBOL/SYMBOOL</th>
<th>VALUE/WAARDE</th>
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<tbody>
<tr>
<td>Avogadro’s constant</td>
<td>$N_A$</td>
<td>$6.02 \times 10^{23}$ mol$^{-1}$</td>
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<tr>
<td>Molar gas constant</td>
<td>$R$</td>
<td>8.31 J·K$^{-1}$·mol$^{-1}$</td>
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<tr>
<td>Standard pressure</td>
<td>$p$</td>
<td>1.013 x $10^5$ Pa</td>
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<tr>
<td>Molar gas volume at STP</td>
<td>$V_m$</td>
<td>22.4 dm$^3$·mol$^{-1}$</td>
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<tr>
<td>Standard temperature</td>
<td>$T$</td>
<td>273 K</td>
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**TABLE 2: FORMULAE**

$$\frac{\dot{V}}{\dot{N}} = \frac{\dot{V}}{\dot{n}} \quad \frac{\dot{V}}{\dot{N}} = \frac{\dot{V}}{\dot{m}}$$

$$T = \frac{1}{2} \frac{\dot{V}}{\dot{N}}$$

$$\frac{\dot{V}}{\dot{m}} = \frac{T}{R}$$

**TABLE 1: PHYSICAL CONSTANTS**

<table>
<thead>
<tr>
<th>Standard temperature</th>
<th>$T$</th>
<th>273 K</th>
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<tr>
<td>Molar gas volume at STP</td>
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TABLE 3: THE PERIODIC TABLE OF ELEMENTS

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</tr>
</tbody>
</table>

Additional Elements:
- Th (232)
- Pa (238)
- U (238)
- Np (93)
- Pu (94)
- Am (95)
- Cm (96)
- Bk (97)
- Cf (98)
- Es (99)
- Fm (100)
- Md (101)
- No (102)
- Lr (103)

Approximate relative atomic mass
Benaderde relatiewe atoommassa

Electronegativity
Elektronegativiteit

Symbol
Simbool
<table>
<thead>
<tr>
<th>Species</th>
<th>Standard Reduction Potential (V)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Li⁺ + e</td>
<td>1.0</td>
</tr>
<tr>
<td>K⁺ + e</td>
<td>1.0</td>
</tr>
<tr>
<td>Cs⁺ + e</td>
<td>1.0</td>
</tr>
<tr>
<td>Sr²⁺ + 2e</td>
<td>0.71</td>
</tr>
<tr>
<td>Ca²⁺ + 2e</td>
<td>0.52</td>
</tr>
<tr>
<td>Fe²⁺ + 2e</td>
<td>0.41</td>
</tr>
<tr>
<td>Co²⁺ + 2e</td>
<td>0.27</td>
</tr>
<tr>
<td>Ni²⁺ + 2e</td>
<td>0.06</td>
</tr>
<tr>
<td>Fe³⁺ + 2e</td>
<td>0.44</td>
</tr>
<tr>
<td>Cd²⁺</td>
<td>0.15</td>
</tr>
<tr>
<td>Pb²⁺</td>
<td>0.06</td>
</tr>
<tr>
<td>Zn²⁺ + 3e</td>
<td>0.74</td>
</tr>
<tr>
<td>Cr³⁺ + 2e</td>
<td>0.91</td>
</tr>
<tr>
<td>H₂O + 2e</td>
<td>0.83</td>
</tr>
<tr>
<td>H₂ + 2e</td>
<td>0.85</td>
</tr>
<tr>
<td>NO + e</td>
<td>1.81</td>
</tr>
<tr>
<td>NO₂⁻ + e</td>
<td>1.23</td>
</tr>
<tr>
<td>NO₃⁻ + 3e</td>
<td>0.96</td>
</tr>
<tr>
<td>H⁺ + e</td>
<td>0.00</td>
</tr>
<tr>
<td>H₂⁻</td>
<td>0.00</td>
</tr>
<tr>
<td>O₂⁻</td>
<td>0.00</td>
</tr>
<tr>
<td>SO₂⁻ + e</td>
<td>0.00</td>
</tr>
<tr>
<td>SO₃⁻ + 2e</td>
<td>0.00</td>
</tr>
<tr>
<td>SO₄²⁻ + 2e</td>
<td>0.00</td>
</tr>
<tr>
<td>Cl⁻ + e</td>
<td>0.00</td>
</tr>
<tr>
<td>Br⁻ + e</td>
<td>0.00</td>
</tr>
<tr>
<td>I⁻ + e</td>
<td>0.00</td>
</tr>
</tbody>
</table>

Increasing oxidising ability / Toenameende oksiderende vermøg

Increasing reducing ability / Toenemende reducerende vermøg

**Table 4A: Standard Reduction Potentials**
### Table 4B: Standard Reduction Potentials

<table>
<thead>
<tr>
<th>( E^\circ (V) )</th>
<th>Half-REACTIONS</th>
</tr>
</thead>
<tbody>
<tr>
<td>+1.00</td>
<td>Fe(^{3+}) + 2e \rightarrow Fe(^{2+})</td>
</tr>
<tr>
<td>+0.80</td>
<td>O(^2-) + 2H(^+) + 2e \rightarrow H(_2)</td>
</tr>
<tr>
<td>+0.68</td>
<td>H(_2)O \rightarrow 2H(^+) + O(^2-) + e</td>
</tr>
<tr>
<td>+0.54</td>
<td>Cu \rightarrow Cu(^{2+}) + 2e</td>
</tr>
<tr>
<td>+0.50</td>
<td>H(_2) \rightarrow 2H(^+) + 2e</td>
</tr>
<tr>
<td>+0.45</td>
<td>S(_2)O(_3(^2-) \rightarrow 2SO(^\cdot) + 2H(^+) + 2e</td>
</tr>
<tr>
<td>+0.40</td>
<td>H(_2)SO(_4) \rightarrow 2H(^+) + SO(_4(^2-) + e</td>
</tr>
<tr>
<td>+0.34</td>
<td>Cu(^{2+}) \rightarrow Cu + 2e</td>
</tr>
<tr>
<td>+0.16</td>
<td>C(_3)H(_6) \rightarrow 2CH(_2) + 2e</td>
</tr>
<tr>
<td>+0.15</td>
<td>Cu(^{2+}) \rightarrow Cu + 2e</td>
</tr>
<tr>
<td>+0.14</td>
<td>H(_2)S \rightarrow 2H(^+) + S(^2-) + 2e</td>
</tr>
</tbody>
</table>

### Increasing Reducing Ability

- Increasing reducing ability indicates a decrease in the standard reduction potential, making it easier for a substance to act as a reducing agent.

### Increasing Oxidizing Ability

- Increasing oxidizing ability indicates an increase in the standard reduction potential, making it easier for a substance to act as an oxidizing agent.
**Atomic Combinations**

**Intramolecular bond**: bond which occurs between atoms within molecules.

**Electronegativity**: a measure of the tendency of an atom to attract a bonding pair of electrons.

**Covalent bond**: a sharing of at least one pair of electrons by two atoms.

- **Non-polar covalent (pure covalent)**: an equal sharing of electrons.
- **Polar covalent**: unequal sharing of electrons leading to a dipole forming (as a result of electronegativity difference).

**Ionic bond**: a transfer of electrons and subsequent electrostatic attraction.

**Metallic bonding**: being between a positive kernel and a sea of delocalized electrons.

**Intermolecular Forces**

- **Ion-dipole forces**: occurs between a polar molecule and an ion.
- **Ion-induced dipole forces**: occurs between an ion and a non-polar molecule.
- **Van der Waals forces**: weak electrostatic forces.
- **Dipole-dipole forces**: occurs between two polar molecules.
- **Dipole-induced dipole forces**: occurs between a polar and a non-polar molecule.
- **Induced-dipole-induced dipole forces** (London Dispersion): occurs between two non-polar molecules.

**Explain hydrogen bonds**: strong intermolecular forces that occur in molecules containing hydrogen bonded to nitrogen, oxygen or fluorine.

**Intermolecular force**: a weak force of attraction between molecules or between atoms of noble gases.

---

**Ideal Gases and Thermal Properties**

- **Standard conditions**: temperature of 0°C (273K) and pressure of 1.01x10^5 Pa (1atm).
- **Boyle’s law**: the volume of an enclosed gas is inversely proportional to its pressure, provided the temperature remains constant.
- **Charles’ law**: the volume of an enclosed gas is directly proportional to the absolute temperature provided the pressure remains constant.
- **Guy-Lussac’s law**: the pressure of an enclosed gas is directly proportional to its temperature on the Kelvin scale, at a constant volume.

**Energy and Chemical Change**

- **Heat of reaction** (ΔH): the net change of chemical potential energy of the system.
- **Exothermic reactions**: reactions which transfer potential energy into thermal energy.
- **Endothermic reactions**: reactions which transfer thermal energy into potential energy.
- **Activation energy**: the minimum energy needed for a reaction to take place.
- **Activated complex**: the unstable transition state from reactants to products.

**Arrhenius theory**

- **Acid**: a substance that produces hydrogen ions (H+)/hydronium ions (H3O+) when it dissolves in water.
- **Base**: a substance that produces hydroxide ions (OH-) when it dissolves in water.

**Lowry-Brønsted theory**

- **Acid**: a proton (H+ ion) donor.
- **Base**: a proton (H+ ion) acceptor.

- **Strong acids**: ionise completely in water to form a high concentration of H3O+ ions. Eg. hydrochloric acid, sulphuric acid, nitric acid.
- **Weak acids**: ionise incompletely in water to form a low concentration of H3O+ ions. Eg. ethanoic acid, oxalic acid.

- **Strong bases**: dissociate completely in water to form a high concentration of OH- ions. Eg. sodium hydroxide, calcium hydroxide.
- **Weak bases**: dissociate/ionise incompletely in water to form a low concentration of OH- ions. Eg. ammonia, calcium carbonate, potassium carbonate.

- **Concentrated acids/bases**: contain a large amount (number of moles) of acid/base in proportion to the volume of water.
- **Dilute acids/bases**: contain a small amount (number of moles) of acid/base in proportion to the volume of water.

---

**Redox Reactions**

- **Oxidation number**: a number assigned to each element in a compound in order to keep track of the movement of electrons.
- **Redox Reaction**: a reaction where oxidation takes place.
- **Reduction**: a reaction where reduction takes place.
- **Oxidising agent**: a substance that is reduced/gains electrons.
- **Reducing agent**: a substance that is oxidised/loses electrons.
- **Anode**: the electrode where oxidation takes place.
- **Cathode**: the electrode where reduction takes place.
- **Electrolyte**: a solution/liquid that conducts electricity through the movement of ions.

---

**Atomic Combinations**

**Intermolecular forces**

- **London dispersion forces**: weak intermolecular forces that occur between non-polar molecules.
- **Hydrogen bonding**: strong intermolecular forces that occur between molecules containing hydrogen bonded to nitrogen, oxygen or fluorine.

---

**Grade 11 Science Essentials**

For more information about Science Clinic’s seminars, classes and resources, visit www.scienceclinic.co.za.
**MOLECULAR STRUCTURE**

Knowing the structure of a molecule enables you to determine its physical and chemical properties. The structure mainly depends on the type of chemical bond (force) that exist between the atoms that the molecule consists of.

### CHEMICAL BONDS

**Chemical bond:** the net electrostatic force that two atoms sharing electrons exert on each other.

Bonding occurs when valence electrons are shared between two atoms or transferred from one atom to another. Valence electrons correspond to the group number of an element on the Periodic table.

**Valence electrons:** electrons in the outermost energy level of an atom that take part in bonding.

The type of bond that forms is dependent on the difference in electronegativity ($\Delta EN$) between the atoms:

- $\Delta EN = 0$: Non-polar (pure) covalent (an equal sharing of electrons)
- $\Delta EN < 1$: Weak polar
- $\Delta EN \leq 2,1$: Polar covalent (unequal sharing of electrons leading to a dipole forming)
- $\Delta EN > 2,1$: Ionic (transfer of electrons and subsequent electrostatic attraction)

**FACTORS INFLUENCING BOND LENGTH**

- **Atom size:** The smaller the atoms, the stronger the bond.
- **Bond order:** The more bonds (double, triple bonds etc.) between the atoms, the stronger the bond.
- **Difference in electronegativity ($\Delta EN$):** The greater the $\Delta EN$, the shorter the bond.

**Bond length:** the average distance between the nuclei of two bonded atoms.

As the atoms get closer, the attractive forces get stronger until the minimum possible potential energy is reached (bond energy). The distance between the nuclei of the atoms at the minimum potential energy is the bond length.

If the two atoms get closer than the bond length, they will be forced apart by the repulsive forces, increasing the potential energy.

### BOND ENERGY

There are various attractive and repulsive forces at play between the two atoms during bonding.

- Attractive forces between the protons of one atom and the electrons of another.
- Attractive forces between the protons and electrons from the same atom.
- A repulsive force between the protons from each atom.
- A repulsive force between the electrons from each atom.

The net electrostatic forces will determine bond strength, which can be quantified as the bond energy. This is the energy required to break the bond, or energy released when bonds are formed.

**Bond energy** is the amount of energy required to break one mol of a specific covalent bond in gaseous phase.

**FACTORS INFLUENCING BOND STRENGTH**

**Bond length**

The shorter the length of the bond, the stronger the bond.

**Atom size**

The smaller the atoms, the stronger the bond.

**Bond order**

The more bonds (double, triple bonds etc.) between the atoms, the stronger the bond.

### BOND LENGTH

The bond length: the average distance between the nuclei of two bonded atoms.

As the atoms get closer, the attractive forces get stronger until the minimum possible potential energy is reached (bond energy). This is the energy required to break the bond, or energy released when bonds are formed.

If the two atoms get closer than the bonding length, they will be forced apart by the repulsive forces, increasing the potential energy.

### A) Covalent Bonding

**Between non-metals**

Covalent bonding is the sharing of at least one pair of electrons by two atoms.

- **Non-polar (pure) covalent**
  - An equal sharing of electrons
  - Eg. $H - P$ bond: $\Delta EN = EN(P) - EN(H) = 0$

- **Weak polar covalent**
  - An unequal sharing of electrons leading to a dipole forming
  - Eg. $H - Br$ bond: $\Delta EN = EN(\text{Br}) - EN(H) = 0,7$

- **Polar covalent**
  - An unequal sharing of electrons leading to a dipole forming
  - Eg. $H - O$ bond: $\Delta EN = EN(O) - EN(H) = 1,4$

### B) Ionic Bonding

**Between metals and non-metals**

Ionic bonding is a transfer of electrons and subsequent electrostatic attraction.

1. Involves a complete transfer of electron(s).
2. Metal atom gives e\(^{-}\) to non-metal.
3. Metal forms a positive cation.
4. Non-metal forms a negative anion.
5. Electrostatic attraction of ions leads to formation of giant crystal lattice.

Ionic Bonding takes place in two steps.

1. Transfer of e\(^{-}\) from cations to anions
2. **Electrostatic attraction**

### C) Metallic Bonding

**Between metals**

Metallic bonding is the bond between the positive metal kernels and the sea of delocalized electrons.

The metal atoms release their valence electrons to surround them. There is a strong but flexible bond between the positive metal kernels and a sea of delocalised electrons.
**LEWIS DIAGRAMS**

Molecules and the atoms that they are made up of are extremely small and cannot be seen with the naked eye. Models are used to visualize the atoms and the chemical bonds between them. One of the ways that we visualize chemical bonds is by means of Lewis diagrams.

Lewis diagrams show the valence electrons around the nucleus, represented by the element symbol.

**EXAMPLE:**
*Give the Lewis-diagrams for the following elements: Li, C, F*

Li: ![Lewis diagram for Li](image)

C: ![Lewis diagram for C](image)

F: ![Lewis diagram for F](image)

**General steps to draw Lewis diagrams**

Lewis diagrams can be drawn for a variety of molecules by using the following general steps:

1. Choose the central atom: the one with the lowest electronegativity (found on the Periodic Table).
2. Determine the total number of valence electrons.
3. Put one shared pair between the central and terminal atoms.
4. Arrange the remaining electrons around the atoms so that 8 electrons are placed around the terminal atoms. (except hydrogen = 2 electrons)
5. If there are not 8 valence electrons around the central atom, move some of the electrons in between the atoms to form multiple bonds.

**EXAMPLE:**
*Give the Lewis diagram for the CO₂ molecule.*

1. C is the central atom as it has an electronegativity of 2.5.
2. CO₂ has 4 + 6 + 6 = 16 valence electrons = 8 electron pairs.
3. O:O
4. 4 electrons have been used: 16 - 4 = 12 electrons are left.
5. Two bonds (a double bond) between each C and O are formed.

**DATIVE BONDING/COORDINATE BONDING**

A coordinate bond is formed when one atom donates both electrons to form a bonding pair. The filled electron orbital (lone pair) of one atom overlaps with the empty orbital of another, e.g. H₂O⁺, NH₄⁺.

**VALENCE SHELL ELECTRON PAIR REPULSION (VSEPR)**

Valence shell electron pair repulsion theory: a model used to predict the shape of a molecule based on the number of electron pairs that surround the central atom.

There are five ideal shapes, where there are no lone pairs on the central atom:

<table>
<thead>
<tr>
<th>Shape name</th>
<th>General formula</th>
<th>Terminal atoms</th>
<th>Lone pairs</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>Linear</td>
<td>*AX₂</td>
<td>2</td>
<td>0</td>
<td>Cl–Be–Cl</td>
</tr>
<tr>
<td>Trigonal planar</td>
<td>*AX₃</td>
<td>3</td>
<td>0</td>
<td>F–B–F</td>
</tr>
<tr>
<td>Tetrahedral</td>
<td>*AX₄</td>
<td>4</td>
<td>0</td>
<td>H–C–H</td>
</tr>
<tr>
<td>Trigonal bipyramidal</td>
<td>*AX₅</td>
<td>5</td>
<td>0</td>
<td>Cl–F–Cl–Cl</td>
</tr>
<tr>
<td>Octahedral</td>
<td>*AX₆</td>
<td>6</td>
<td>0</td>
<td>F–F</td>
</tr>
</tbody>
</table>

*A is the central atom and X represents the terminal atoms.

Molecules that have lone pairs will have a slightly different shape, since the lone pairs exert repulsive forces on the other bonding pairs, pushing them down.

**BOND VS MOLECULAR POLARITY**

**Bond polarity**

Bond polarity is determined by the electronegativity of the 2 atoms that form the bond.

**Electronegativity** is the tendency of an atom to attract a shared electron pair during bonding.

If one atom has a greater electronegativity than another, the electrons will be pulled more towards one atom than another. As such a shift of electrons creates negative and positive charge distributions inside the molecule.

**Polar bond**: an unequal distribution of electrons between two atoms during bonding.

Non-polar bond: an equal distribution of electrons between two atoms during bonding.

**Molecular polarity**

A molecule can have polar bonds, but not be polar as a whole. Molecular polarity is determined by the shape of the molecule.

**Polar molecule**: a molecule over which charge is distributed unevenly.

Non-polar molecule: a molecule over which the charge is evenly distributed.

Lone pairs + EN (O): O is slightly negative.

Lower EN (H): H is slightly positive.

Uneven charge distribution: polar molecule.

EN (F): F is slightly negative.

EN (B): B is slightly positive.

F atoms evenly distributed: non-polar molecule.
INTERMOLECULAR FORCES (IMF)

Intermolecular forces are weak forces of attraction between molecules or between atoms of noble gases.

IMF vs Intramolecular bonds

Intermolecular forces are not the same as intramolecular bonds.

Intramolecular bonds exist between atoms in a molecule.

Intermolecular forces exist between molecules or between atoms of noble gases.

AMAZING WATER MOLECULE

• Non-symmetrical polar molecule.

• Intramolecular covalent bonds.

• Bent/angular molecular shape.

• Contains hydrogen bonding between molecules. These strong IMF need a large amount of energy to overcome.

• Water vapour acts as a greenhouse gas by absorbing the sun’s radiation.

• The ocean acts as a heat reservoir and regulates the temperature of earth.

PROPERTIES OF WATER

Specific heat capacity

The amount of energy needed to raise the temperature of 1 kg of water by 1 °C. Water has a high specific heat capacity.

A high heat capacity requires more energy for the temperature to change. This helps with heat regulation on the earth.

Melting and boiling points

Strong hydrogen bonding IMF’s require large amounts of energy to overcome intermolecular forces.

Melting point: 0 °C

Boiling point: 100 °C at atmospheric pressure.

Density

As water freezes the molecules orientate themselves so that the volume expand and the density decreases. Ice forms a crystal lattice.

More dense

Less dense

EXAMPLE:

Complete the table below by identifying the dominant intermolecular force present in each example.

<table>
<thead>
<tr>
<th>Substance</th>
<th>Intermolecular force</th>
</tr>
</thead>
<tbody>
<tr>
<td>CO₂</td>
<td>London dispersion</td>
</tr>
<tr>
<td>CO</td>
<td>Dipole-dipole</td>
</tr>
<tr>
<td>NH₃</td>
<td>Hydrogen bonding</td>
</tr>
<tr>
<td>NaCl in water</td>
<td>Ion-dipole</td>
</tr>
<tr>
<td>HF</td>
<td>Hydrogen bonding</td>
</tr>
<tr>
<td>NO and CH₄ gas mix</td>
<td>Dipole-induced dipole</td>
</tr>
<tr>
<td>HCl</td>
<td>Dipole-dipole</td>
</tr>
<tr>
<td>KMnO₄ in CHCl₃</td>
<td>Ion-dipole</td>
</tr>
</tbody>
</table>
IMF AND PHYSICAL PROPERTIES

**Phase of molecules**
IMF between molecules causes them to be attracted to one another.

**Stronger IMF = Stronger attraction**
At the same temperature:
Solid: Strong IMF
Liquid: Intermediate IMF
Gas: Weak IMF

**Density**
Density is the number of particles per unit volume.
Increased IMF increases the attractive force between the particles, causing the particles to become more tightly packed. The arrangement of the particles determine the number of particles per volume, thereby determining the density.

**Stronger IMF = Particles closer together = Greater density**

**Phase change**
Melting: Melting points increase with increase in IMF.
Boiling: Boiling points increase with increase in IMF.

**Viscosity**
Viscosity is a liquids resistance to flow.
Stronger IMF will cause the molecules in the liquid to hold together more strongly, causing the viscosity to increase.

**Stronger IMF = Greater viscosity**

**Capillarity**
Capillarity is the tendency of liquid to rise in a tube.
Capillary action occurs when the adhesive forces are stronger than the cohesive forces.

**Stronger IMF between liquid and tube = Greater capillary action**

**Surface tension**
Surface tension is the cohesive forces between molecules on the surface of a liquid.
Molecules at the surface of a liquid experience only inward intermolecular forces.

**Stronger IMF = Greater surface tension**

**Thermal expansion**
Thermal expansion is the expanding of a material when heated.
As the temperature increases, the kinetic energy of the molecules increase. This additional energy can overcome the IMF and molecules move further apart.

**Stronger IMF = molecules held together strongly = Lower thermal expansion**

**Solubility**
Solubility is the ability of a solute to dissolve in a solvent.
A solute will dissolve in a solvent with similar IMF’s. “like dissolves like”.

**Polar and ionic solute will dissolve in a polar solvent**
A non-polar solute will dissolve in a non-polar solvent.

**Adhesion**
Adhesion is the intermolecular attraction of molecules to molecules of a different substance.
The strength of adhesive forces are dependent on the two specific substances interacting with each other.

**Cohesion**
Cohesion is the intermolecular attraction between molecules of the same substance.

**Stronger IMF = Greater cohesive forces**

**Size of molecules**
Larger molecule = Larger electron cloud = Stronger IMF

**Density**
Density is the number of particles per unit volume.
Increased IMF increases the attractive force between the particles, causing the particles to become more tightly packed. The arrangement of the particles determine the number of particles per volume, thereby determining the density.

**Stronger IMF = Particles closer together = Greater density**

**Viscosity**
Viscosity is a liquids resistance to flow.
Stronger IMF will cause the molecules in the liquid to hold together more strongly, causing the viscosity to increase.

**Stronger IMF = Greater viscosity**

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Surface tension is the cohesive forces between molecules on the surface of a liquid.
Molecules at the surface of a liquid experience only inward intermolecular forces.

**Stronger IMF = Greater surface tension**

**Solubility**
Solubility is the ability of a solute to dissolve in a solvent.
A solute will dissolve in a solvent with similar IMF’s. “like dissolves like”.

**Polar and ionic solute will dissolve in a polar solvent**
A non-polar solute will dissolve in a non-polar solvent.

**HF contains hydrogen bonding→strongest IMF→highest boiling point. HCl, HBr and HI all contain dipole-dipole forces and therefore the boiling point increases with increasing molecular mass.**

**Evaporation:** Evaporation decreases with increase in IMF. Higher vapor pressures correspond to weaker IMF.
IDEAL GASES

Properties of gases according to kinetic molecular theory

- Particles are far apart, large spaces between them.
- Particles move with Brownian motion, Fast, random motion.
- Gases undergo diffusion, Particles spread out to fill the container.
- The forces between particles are negligibly small.
- Particles collide elastically with other particles and the walls of the container.

Three measurable variables:
- Temperature (T) - a measure of average kinetic energy.
- Volume (V) - the space occupied by the gas in the container.
- Pressure (p) - the outward force per unit area applied by gas when the gas particles collide with the container walls.

**Standard Temperature and Pressure or STP**

Standard Temperature = 0 °C or 273 K
Standard Pressure = 101.3 kPa (1.013 x 10^5 Pa)

IDEAL GASES

Ideal gases are theoretical gases that obey the gas laws under all conditions of temperature and pressure.

- The particles of a gas are all identical and in constant motion.
- The volume of the gas is due to the motion of the particles as the particles have no volume themselves.
- The intermolecular forces between gas particles are negligible.
- All collisions are perfectly elastic.

**DEVIATION FROM IDEAL GAS**

At LOW temperatures:
- Gas particles move much slower, decreasing the number of collisions between the particles and the container walls. Pressure of the gas is lower than expected.
- Attractive forces are stronger due to the particles being closer together and moving slower. The gas is more likely to become a liquid and the volume is greater than expected.

At HIGH pressure:
- The volume of the gas particles becomes significant in relation to the container.
- Increases in the intermolecular forces between the particles due to proximity, gas will liquefy and the volume is larger than expected.

**BOYLE’S LAW**

\[ p_1V_1 = p_2V_2 \]

The volume of an enclosed gas is inversely proportional to the pressure of the gas if the temperature remains constant.

**CHARLES’ LAW**

\[ \frac{V_1}{T_1} = \frac{V_2}{T_2} \]

The volume of an enclosed gas is directly proportional to the temperature (in Kelvin) of the gas provided the pressure remains constant.

**GAY LUSSAC’S LAW**

\[ \frac{p_1}{T_1} = \frac{p_2}{T_2} \]

The pressure of an enclosed gas is directly proportional to the absolute temperature if the volume of the gas remains constant.

**IDEAL GAS EQUATION**

\[ pV = nRT \]

**Example:**

A certain sample of gas has a volume of 125 cm³ at standard temperature and pressure. If the pressure of the gas increases to 25 °C, calculate the new volume of the gas.

\[ \frac{V_1}{T_1} = \frac{V_2}{T_2} \]

\[ \frac{125}{273} = \frac{V_2}{298} \]

\[ V_2 = 136.45 \text{ cm}^3 \]

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**Quantitative Aspects of Chemical Change**

### The Mole

Atoms, molecules and ions are too small to count, and there are so many particles in even the smallest sample of a substance.

There are more particles of water in a teaspoon than there are teaspoons of water in all the oceans.

Rather than dealing with the particles individually, we deal with a special number of particles.

The **mole** is a name for a special number. Many numbers have names, such as:

- 2 = pair
- 3 = hat-trick
- 12 = dozen

A mole of particles is an amount of \(6.02 \times 10^{23}\) particles. \(6.02 \times 10^{23}\) is known as Avogadro's number, \(N_A\).

Avogadro's number (\(N_A\)) is too big to imagine.

\[6.02 \times 10^{23}\text{ is known as Avogadro's number, } N_A\]

This many grains of sand, piled on the surface of the earth, would come almost reach the moon.

The mole is defined as the number of particles or atoms in 12.0 g of Carbon -12.

### Molar Mass

Particles are too small to weigh individually.

**Molar mass** (\(M\)) is defined as the mass of one mole of particles (atoms, molecules or formula units) and is measured in the unit g mol\(^{-1}\).

\[
n = \frac{m}{M} \tag{1}\]

\(n\) is number of mole (mol), \(m\) is mass of substance (g), and \(M\) is molar mass (g mol\(^{-1}\)).

Relative atomic mass (\(A_r\)) is the average mass of an atom compared to the mass of a Carbon 12 atom. It is measured in atomic mass units (amu).

Molar mass (\(M\)) of an element is equal to the magnitude of relative atomic mass (\(A_r\)) in amu. This is found on the periodic table. See the table below for other substances:

<table>
<thead>
<tr>
<th>Type of substance</th>
<th>Particles</th>
<th>Example</th>
<th>Formula</th>
<th>Molar mass g mol(^{-1})</th>
</tr>
</thead>
<tbody>
<tr>
<td>Element</td>
<td>Atoms</td>
<td>Neon</td>
<td>Ne</td>
<td>20</td>
</tr>
<tr>
<td>Covalent compound</td>
<td>Molecules</td>
<td>Carbon dioxide</td>
<td>CO(_2)</td>
<td>12 + 32 = 44</td>
</tr>
<tr>
<td>Ionic compound</td>
<td>Ions (formula units)</td>
<td>Salt</td>
<td>NaCl</td>
<td>23 + 35.5 = 58.5</td>
</tr>
<tr>
<td>Metallic compound</td>
<td>Positive kernels and delocalized electrons</td>
<td>Gold</td>
<td>Au</td>
<td>197</td>
</tr>
</tbody>
</table>

**EXAMPLE:**

What is the relative formula mass of Calcium sulphate (CaSO\(_4\))?

\[M_r(CaSO_4) = M_r(Ca) + M_r(S) + (4 \times M_r(O))\]

\[= 40 + 32 + (4 \times 16)\]

\[= 136 \text{ (no unit)}\]

**EXAMPLE:**

What is the molar mass of Calcium sulphate (CaSO\(_4\))? 136 g mol\(^{-1}\)

### Percentage Composition

**Percentage composition of element = \(\frac{\text{molar mass of element}}{M_r \text{ of compound}} \times 100\)**

Consider these iron ores: haematite and magnetite – which contains more iron by mass?

<table>
<thead>
<tr>
<th>Ore</th>
<th>Haematite</th>
<th>Magnetite</th>
</tr>
</thead>
<tbody>
<tr>
<td>Formula</td>
<td>Fe(_2)O(_3)</td>
<td>Fe(_3)O(_4)</td>
</tr>
<tr>
<td>Relative molecular mass</td>
<td>((2 \times 56) + (3 \times 16) = 160)</td>
<td>((3 \times 56) + (4 \times 16) = 232)</td>
</tr>
<tr>
<td>% iron by mass</td>
<td>([(2 \times 56) / 160] \times 100 = 70%)</td>
<td>([(3 \times 56) / 232] \times 100 = 72%)</td>
</tr>
</tbody>
</table>

\[\therefore\text{ magnetite contains more iron}\]

### Different Types of Chemical Formulae

Consider the substance ethane.

It also can be represented using a formula. There are three types of formulas we use:

- **Molecular formula**: Actual number of each atom. Eg. C\(_2\)H\(_6\)
- **Empirical formula**: Simplest whole number ratio of the atoms. Eg. C\(_2\)H\(_6\) → CH\(_3\)
- **Structural formula**: Shows how the atoms are joined.

\[\begin{align*}
\text{Molecular formula} & \quad \text{Actual number of each atom.} \\
\text{Empirical formula} & \quad \text{Simplest whole number ratio of the atoms.} \\
\text{Structural formula} & \quad \text{Shows how the atoms are joined.}
\end{align*}\]

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Quantitative Aspects of Chemical Change

Concentrations of Solutions

Solutions are homogeneous (uniform) mixtures of two or more substances. A solution is formed when a solute dissolves in a solvent. The solvent and solute can be a gas, liquid, or solid. The most common solvent is liquid water. This is called an aqueous solution.

<table>
<thead>
<tr>
<th>Solution</th>
<th>Solute</th>
<th>Solvent</th>
</tr>
</thead>
<tbody>
<tr>
<td>salt water</td>
<td>Salt</td>
<td>water</td>
</tr>
<tr>
<td>Soda water</td>
<td>Carbon dioxide</td>
<td>water</td>
</tr>
</tbody>
</table>

Concentration

The concentration of a solution is the number of moles of solute per unit volume of solution.

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

where:
- \( C \) is the concentration (mol dm\(^{-3} \))
- \( n \) is the number of moles (mol)
- \( V \) is the volume (dm\(^3 \))

can also be calculated with

\[ c = \frac{m}{MV} \]

i.e. The number moles of solute per 1 dm\(^3 \) of solution i.e. mol dm\(^{-3} \). If a solution of potassium permanganate \( \text{K}_\text{MnO}_4 \) has a concentration of 2 mol dm\(^{-3} \) it means that for every 1 dm\(^3 \) of solution, there are 2 moles of \( \text{K}_\text{MnO}_4 \) dissolved in the solvent.

**EXAMPLE:**

A solution contains 10 g of sodium hydroxide, \( \text{NaOH} \), in 200 cm\(^3 \) of solution. Calculate the concentration of the solution.

\[
\begin{align*}
    n(\text{NaOH}) &= \frac{m}{M} \\
    &= \frac{10}{23 + 16 + 1} \\
    &= 0.25 \text{ mol}
\end{align*}
\]

\[
\begin{align*}
    V &= 200 \text{ cm}^3 = 0.2 \text{ dm}^3 \\
    c(\text{NaOH}) &= \frac{n}{V} \\
    &= \frac{0.25}{0.2} \\
    &= 1.25 \text{ mol dm}^{-3}
\end{align*}
\]

Molar Volumes of Gases

If different gases have the same volume under the same conditions of temperature and pressure, they will have the same number of molecules.

The molar volume of a gas, \( V_m \), is the volume occupied by one mole of the gas.

\( V_m \) for all gases is 22.4 dm\(^3\) mol\(^{-1} \).

**Standard Temperature and Pressure (STP)** is 273 K (0°C) and 1.01x10\(^5 \) Pa.

This also means that for reactions at constant temperature and pressure, gas volumes will react in the same ratio as the molar ratio.

\[
\begin{align*}
    \text{N}_2 + 2\text{O}_2 &\rightarrow 2\text{NO}_2 \\
    1 \text{ mol} + 2 \text{ mol} &\rightarrow 2 \text{ mol} \\
    1 \text{ dm}^3 + 2 \text{ dm}^3 &\rightarrow 2 \text{ dm}^3
\end{align*}
\]

Concentrations of solutions

**EXAMPLE:**

Calculate the mass of solute in 600 cm\(^3 \) of 1.5 mol dm\(^{-3} \) sodium chloride solution.

\[
\begin{align*}
    V &= 600 \text{ cm}^3 = 0.6 \text{ dm}^3 \\
    M(\text{NaCl}) &= 23 + 35.5 \\
    &= 58.5 \text{ g mol}^{-1}
\end{align*}
\]

\[
\begin{align*}
    n &= \frac{V}{V_M} \\
    &= \frac{0.6}{22.4} \\
    &= 0.01 \text{ mol}
\end{align*}
\]

**EXAMPLE:**

A gas jar with a volume of 224 cm\(^3 \) is full of chlorine gas, at STP. How many moles of chlorine gas are there in the gas jar?

\[
\begin{align*}
    n &= \frac{V}{V_M} \\
    &= \frac{0.224}{22.4} \\
    &= 0.01 \text{ mol}
\end{align*}
\]

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Quantitative aspects of chemical change

Water of crystallization
Some ionic crystals trap a certain number of water molecules between the ions as they are forming. These water molecules are known as "Water of crystallization".

Eg. Hydrated copper sulphate:
CuSO₄·5H₂O has 5 water molecules per formula unit. When the hydrated salt crystals are heated, the water molecules evaporate off, leaving the anhydrous salt behind.

To calculate the number of moles of water of crystallization:
1. Calculate the mass of water that evaporated off.
2. Calculate the moles of each substance.
3. Determine the ratio of water to anhydrous salt.
4. Write the formula for the hydrated salt.

EXAMPLE:
13.2 g of a sample of zinc sulphate, ZnSO₄·xH₂O, was heated in a crucible. Calculate the number of moles of water of crystallisation if 7.4 g of solid remained.

Empirical formula: ZnSO₄·7H₂O

Calculating the Empirical Formula from Percentage Composition
The empirical formula of a compound can also be found from its percentage composition. We assume that 100 g of the compound is analysed, then each percentage gives the mass of the element in grams in 100 g of the compound.

Calculating the Empirical Formula of a Compound from Mass
Empirical formula is the chemical formula of a compound that shows the smallest whole number ratio of the atoms.

To calculate the empirical formula of a compound from mass:
1. Determine the mass of the elements.
2. Determine mol of each substance.
3. Simplify the atomic ratio.

EXAMPLE:
In a combustion reaction 0.48 g of magnesium ribbon is burnt. The amount of magnesium oxide produced is 0.80 g.

Calculate the empirical formula for magnesium oxide.
Empirical formula to Molecular Formula

The empirical formula is the simplest whole number ratio of atoms in a molecule. The molecular formula is the actual ratio of the atoms in a molecule.

The molecular formula can be calculated from the empirical formula and the relative molecular mass.

**STEPS TO DETERMINE MOLECULAR FORMULA:**

1. Determine the empirical formula (if not given).
2. Determine the molar mass of the empirical formula.
3. Determine the molar ratio of the reactants to products.
4. Determine the molar ratio between molecular formula and empirical formula.
5. Multiply the ratio into the empirical formula

**Example:**

Determine the molecular formula of butene. The molecular mass of butene is 56 g mol⁻¹. Determine the molecular formula of butene.

1. Empirical formula given: CH₂
2. \( M(CH_2) = 12 + 1 + 1 = 14 \text{ g mol}^{-1} \)
3. \( \text{ratio number} = \frac{\text{molecular formula mass}}{\text{empirical formula mass}} = \frac{56}{14} = 4 \)
4. \( CH_2 \times 4 = C_4H_8 \)

**Quantitative Aspects - Reactions**

**Approach to reaction stoichiometry**

1. Write a balanced chemical equation.
2. Change the 'given' amount into mole (use limiting reactant if applicable).
3. Determine the number of mole of the 'asked' substance using the mole ratio.
4. Determine the 'asked' amount from the number of mole.

**Example:**

To obtain iron, iron (III) oxide reacts with carbon monoxide according to the following word equation:

- Iron (III) oxide + carbon monoxide → iron + carbon dioxide
- What mass of iron is produced from 48 g of iron (III) oxide?

1. \( \text{Fe}_2\text{O}_3(s) + 3\text{CO}(g) \rightarrow 2\text{Fe}(s) + 3\text{CO}_2(g) \)
2. \( n(\text{Fe}_2\text{O}_3) = \frac{m}{M} = \frac{48}{160} = 0.3 \text{ mol} \)
3. \( \text{Fe}_2\text{O}_3 : \text{Fe} = 1 : 2 \)
   - 0.3 mol : 0.6 mol
4. \( m(\text{Fe}) = nM = (0.6)(56) = 33.6 \text{ g} \)

**Balanced chemical equations**

Eg. \( \text{N}_2(g) + 3\text{H}_2(g) \rightarrow 2\text{NH}_3(g) \)

The balanced chemical equation is the "recipe". It tells us:

- The reactants used i.e. Nitrogen and Hydrogen
- The products produced, i.e. ammonia
- The states of the reactants and products (all gases here)
- The molar ratio of the reactants to products.

1 mole of nitrogen reacts with 3 molecules of hydrogen to form 2 molecules of ammonia.

1 mole of nitrogen reacts with 3 moles of hydrogen to form 2 moles of ammonia.

**Limiting Reactants**

In a reaction between two substances, one reactant is likely to be used up completely before the other. This limits the amount of product formed.

Consider the reaction between magnesium and dilute sulphuric acid. The balanced chemical equation is

\( \text{Mg}(s) + \text{H}_2\text{SO}_4(aq) \rightarrow \text{MgSO}_4(aq) + \text{H}_2(g) \)

This means that 1 mole of magnesium reacts with 1 mole of sulphuric acid. Both reactants will be completely used up by the time the reaction stops.

What happens if 1 mole of magnesium reacts with 2 mole of sulphuric acid? There is now insufficient magnesium to react with all of the sulphuric acid. 1 mole of sulphuric acid is left after the reaction.

All of the magnesium is used up, We say the magnesium is the limiting reactant. Some sulphuric acid is left after the reaction. We say the sulphuric acid is in excess.

The amount of limiting reactant will determine:

- The amount of product formed.
- The amount of other (excess) reactants used.

**Determining limiting reactants**

1. Calculate the number of moles of each reactant.
2. Determine the ratio between reactants.
3. Determine limiting reactant using the ratios.

**NOTE:**

If one reactant is in excess, it means that there is more than enough of it. If there are only 2 reactants and one is in excess, it means that the other is the limiting reactant.
EXAMPLE:
A 8,4 g sample of nitrogen reacts with 1,5 g hydrogen. The balanced equation is:

\[ \text{N}_2(g) + 3 \text{H}_2(g) \rightarrow 2 \text{NH}_3(g) \]

Determine (a) which reactant is the limiting reactant, and (b) the mass of ammonia that can be produced.

(a) 1. \( n(\text{N}_2) = \frac{m}{M} = \frac{8.4}{28} = 0.3 \text{ mol} \)

2. \( n(\text{H}_2) = \frac{m}{M} = \frac{1.5}{2} = 0.75 \text{ mol} \)

Hence hydrogen is the limiting reactant.

(b) Because the hydrogen is the limiting reactant, it will determine the mass of ammonia produced:

\[ \begin{align*}
\text{H}_2 & : \text{NH}_3 \\
3 & : 2 \\
0.75 \text{ mol} & : 0.5 \text{ mol} \\
\text{m(} \text{NH}_3 \text{)} & = \frac{n \text{M}}{1} = (0.5)(17) \\
& = 8.5 \text{ g}
\end{align*} \]

Percentage purity = \( \frac{\text{Mass of pure substance}}{\text{Mass of impure substance}} \times 100 \)

\[ \begin{align*}
\text{Percentage purity} & = \frac{8.5}{8.75} \times 100 \\
& \approx 97.5 \%
\end{align*} \]

Steps to determine the percentage purity:
1. Determine moles of products.
2. From the balanced formula, determine the ratio between reactants and products.
3. Using the ratio, determine the number of moles of reactants.
4. Determine the mass of pure reactant.
5. Calculate the percentage purity of the sample.

EXAMPLE:
An impure sample of calcium carbonate, \( \text{CaCO}_3 \) contains calcium sulphate, \( \text{CaSO}_4 \), as an impurity. When excess hydrochloric acid was added to 6 g of the sample, 1200 cm\(^3\) of gas was produced (measured at STP). Calculate the percentage purity of the calcium carbonate sample. The equation for the reaction is:

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

1. \( n(\text{CO}_2) = \frac{V}{V_M} = \frac{1.2}{22.4} = 0.054 \text{ mol} \)

2. \( \frac{\text{CaCO}_3}{\text{CO}_2} = \frac{1}{1} = 0.054 \)

\( \therefore \) 0.054 mol \( \text{CaCO}_3 \) reacted

3. \( m(\text{CaCO}_3) = n \text{M} = (0.054)(40 + 12 + 16 + 16) = 5.4 \text{ g} \)

4. Percentage purity = \( \frac{\text{Mass of pure substance}}{\text{Mass of impure substance}} \times 100 \)

\[ \begin{align*}
\text{Percentage purity} & = \frac{5.4}{6.0} \times 100 \\
& \approx 90 \%
\end{align*} \]

Steps to determine the percentage yield:
1. Determine moles of reactants.
2. From the balanced formula, determine the ratio between reactants and products.
3. Using the ratio, determine the number of moles of products.
4. Determine the theoretical mass of product.
5. Calculate the percentage yield.

EXAMPLE:
128 g of sulphur dioxide, \( \text{SO}_2 \), was reacted with oxygen to produce sulphur trioxide, \( \text{SO}_3 \). The equation for the reaction is:

\[ 2\text{SO}_2(g) + \text{O}_2(g) \rightarrow 2\text{SO}_3(g) \]

140 g of \( \text{SO}_3 \) was produced in the reaction. Calculate the percentage yield of the reaction.

1. \( n(\text{SO}_3) = \frac{m}{M} = \frac{128}{64} = 2 \text{ mol} \)

2. \( \frac{\text{SO}_2}{\text{SO}_3} = \frac{2}{1} \)

3. \( \frac{\text{SO}_2}{\text{SO}_3} = \frac{2}{1} = 2 \text{ mol SO}_3 \)

4. \( m(\text{SO}_3) = n \text{M} = (2)(32 + 16 + 16 + 16) = 160 \text{ g} \)

5. Percentage yield = \( \frac{\text{Mass of product produced}}{\text{Maximum theoretical mass of product}} \times 100 \)

\[ \begin{align*}
\text{Percentage yield} & = \frac{140}{160} \times 100 \\
& = 87.5 \%
\end{align*} \]

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ENERGY CHANGE
The energy change that takes place occurs because of bonds being broken and new bonds being formed.

When bonds are broken, energy is absorbed from the environment.
When bonds are formed, energy is released into the environment.

The net energy change will determine if the reaction is endothermic or exothermic.

ACTIVATION ENERGY
In order to start a reaction, energy first needs to be absorbed to break the bonds. This energy is known as the activation energy - the minimum energy required to start a chemical reaction.

Once the bonds have been broken, the atoms in the chemical system form an activated complex - a temporary transition state between the reactants and the products.

ENTHALPY AND ENTHALPY CHANGE
Enthalpy (H) is the total amount of stored chemical energy (potential energy) of the reactants and the products. During chemical reactions, energy can be exchanged between the chemical system and the environment, resulting in a change in enthalpy. This change in enthalpy, \( \Delta H \), represents the heat of the reaction measured in kJ·mol\(^{-1}\).

The heat of reaction (\( \Delta H \)) is the net change of chemical potential energy of the system.

CHEMICAL SYSTEM AND THE ENVIRONMENT
The chemical system is the reactant and product molecules.

The environment is the surroundings of the chemical system, including the container in which the reaction takes place, or the water in which the molecules are dissolved.

<table>
<thead>
<tr>
<th>ENDOThERMIC</th>
<th>EXOTHERMIC</th>
</tr>
</thead>
<tbody>
<tr>
<td>Definition: A reaction that absorbs heat energy from the surroundings</td>
<td>Definition: A reaction that releases heat energy into the environment</td>
</tr>
<tr>
<td>More energy absorbed than released</td>
<td>More energy released than absorbed</td>
</tr>
<tr>
<td>Net energy change is energy absorbed from the environment</td>
<td>Net energy change is energy released into the environment</td>
</tr>
<tr>
<td>The chemical system's energy increases (( \Delta H &gt; 0 ))</td>
<td>The chemical system's energy decreases (( \Delta H &lt; 0 ))</td>
</tr>
<tr>
<td>The environment's energy decreases</td>
<td>The environment's energy increases</td>
</tr>
<tr>
<td>Temperature of the environment decreases (test tube gets colder)</td>
<td>Temperature of the environment increases (test tube gets hotter)</td>
</tr>
</tbody>
</table>

CATALYST
In order for a reaction to occur, enough energy has to be provided (activation energy) for particles to collide effectively.

The amount of required energy can be decreased by using a catalyst. A catalyst is a chemical substances that lowers the activation energy required without taking part in the reaction. By lowering the activation energy, the rate of the reaction can also be increased.

A catalyst is a substance that increases the rate of the reaction but remains unchanged at the end of the reaction.

IMPORTANT REACTIONS

ENDOTHERMIC
Photosynthesis
\[ 6\text{CO}_2 + 6\text{H}_2\text{O} \xrightarrow{\text{light}} 6\text{C}_6\text{H}_12\text{O}_6 + 6\text{O}_2 ; \Delta H > 0 \]

EXOTHERMIC
Cellular respiration
\[ \text{C}_6\text{H}_12\text{O}_6 + 6\text{O}_2 \rightarrow 6\text{CO}_2 + 6\text{H}_2\text{O} ; \Delta H < 0 \]

Combustion
\[ \text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O} ; \Delta H < 0 \]
\[ \text{C}_2\text{H}_5\text{OH} + 3\text{O}_2 \rightarrow 2\text{CO}_2 + 3\text{H}_2\text{O} ; \Delta H < 0 \]
**ACID/BASE DEFINITIONS**

**Arrhenius**
An acid is a substance that produces hydrogen ions (H\(^+\)) / hydronium ions (H\(_3\)O\(^+\)) when dissolved in water.

A base is a substance that produces hydroxide ions (OH\(^-\)) when dissolved in water.

**Bronsted-Lowry**
An acid is a proton (H\(^+\)) donor.
A base is a proton (H\(^+\)) acceptor.

**ACID PROTICITY**

Some acids are able to donate more than one proton. The number of protons that an acid can donate is referred to as the acid proticity.

1 proton- monoprotic
2 protons- diprotic
3 protons- triprotic

**CONJUGATE ACID-BASE PAIRS**

An acid forms a conjugate base when it donates a proton.

acid = conjugate base + H\(^+\)

A base forms a conjugate acid when it accepts a proton.

base + H\(^+\) = conjugate acid

Conjugate acid-base pairs are compounds that differ by the presence of one proton, or H\(^+\).

**EXAMPLE:**

Identify the conjugate acid-base pair in the following example:

\[
\text{HNO}_3(\text{aq}) + \text{H}_2\text{O}(\ell) \rightarrow \text{NO}_3^- (\text{aq}) + \text{H}_2\text{O}^+ (\text{aq})
\]

**STRONG VS WEAK ACIDS AND BASES**

The strength of an acid/base refers to the ability of the substance to ionise or dissociate.

**ACIDS**

A strong acid will ionise completely in water.

\[
\text{HCl} (\text{g}) + \text{H}_2\text{O}(\ell) \rightarrow \text{H}_3\text{O}^+ (\text{aq}) + \text{Cl}^- (\text{aq})
\]

(stong acid → weak conjugate base)

A weak acid will only partially ionise in water.

\[
2\text{H}_2\text{CO}_3(\ell) + \text{H}_2\text{O}(\ell) \rightarrow 2\text{H}_2\text{O}^+ (\text{aq}) + \text{CO}_2\text{=}^- (\text{aq}) + \text{H}_2\text{CO}_3 (\text{aq})
\]

(weak acid → strong conjugate base)

**BASES**

A strong base will dissociate completely in water.

\[
\text{NaOH} (\text{s}) \rightarrow \text{Na}^+ (\text{aq}) + \text{OH}^- (\text{aq})
\]

(strong base → weak conjugate acid)

A weak base will dissociate only partially in water.

\[
\text{Mg(OH)}_2(\text{s}) \rightarrow \text{Mg}^2+ (\text{aq}) + 2\text{OH}^- (\text{aq}) + \text{Mg(OH)}_2 (\text{aq})
\]

(weak base → strong conjugate acid)

NH\(_3\) is an exception, it ionises.

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

**AMPHOLYTE/ AMPHOTERIC SUBSTANCES**

Ampholyte- A substance that can act as both an acid or a base.

Amphoteric/amphiprotic substances can therefore either donate or accept protons. Common ampholytes include H\(_2\)O, HCO\(_3^-\) and HSO\(_4^-\).

H\(_2\)SO\(_4\) as an ampholyte:

Acid: HSO\(_4^-\) + H\(_2\)O \rightarrow SO\(_4^{2-}\) + H\(_2\)O

Base: HSO\(_4^-\) + H\(_2\)O \rightarrow H\(_2\)SO\(_4\) + OH\(^-\)

**INFLUENCE OF ACID/BASE STRENGTH**

**Reaction rate**

Reaction rates increase as the strength of the acid/base increases.

**Conductivity**

Conductivity increases as the strength of the acid/base increases.

**COMMON ACIDS**

<table>
<thead>
<tr>
<th>Acid</th>
<th>Common Bases</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydrochloric acid (HCl)</td>
<td>Sodium Hydroxide (NaOH)</td>
</tr>
<tr>
<td>Nitric acid (HNO(_3))</td>
<td>Potassium hydroxide (KOH)</td>
</tr>
<tr>
<td>Sulfuric acid (H(_2)SO(_4))</td>
<td>Sodium hydrogen carbonate (NaHCO(_3))</td>
</tr>
<tr>
<td>Hydrofluoric acid (HF)</td>
<td>Calcium carbonate (CaCO(_3))</td>
</tr>
<tr>
<td>Sulfurous acid (H(_2)SO(_3))</td>
<td>Sodium carbonate (Na(_2)CO(_3))</td>
</tr>
<tr>
<td>Carbonic acid (H(_2)CO(_3))</td>
<td>Ammonia (NH(_3))</td>
</tr>
<tr>
<td>Acetic acid / ethanoic acid (CH(_3)COOH)</td>
<td></td>
</tr>
<tr>
<td>Phosphoric acid (H(_3)PO(_4))</td>
<td></td>
</tr>
</tbody>
</table>

**THE pH SCALE**

The pH of a solution is a number that represents the acidity or alkalinity of a substance.

The stronger the concentration of H\(^+\) ions in solution, the more acidic the solution and the lower the pH. The lower the concentration of H\(^+\) in solution, the more alkaline the solution and the higher the pH.

The pH scale is a range from 0 to 14.

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ACIDS AND BASES

PROTOTYPICAL REACTIONS

Prototypical reactions are reactions during which protons (H+) are transferred.

\[
\text{acid} + \text{metal} \rightarrow \text{salt} + \text{H}_2
\]

\[
\text{HA} + \text{M}^+ \rightarrow \text{M}^+ + \text{H}_2
\]

\[
2\text{HNO}_3(aq) + 2\text{Na}(s) \rightarrow 2\text{NaNO}_3(aq) + \text{H}_2(g)
\]

\[
\text{acid} + \text{metal hydroxide (base)} \rightarrow \text{salt} + \text{H}_2\text{O}
\]

\[
\text{HA} + \text{MOH} \rightarrow \text{M}^+ + \text{H}_2\text{O}
\]

\[
\text{acid} + \text{metal oxide} \rightarrow \text{salt} + \text{H}_2\text{O}
\]

\[
\text{HA} + \text{MO} \rightarrow \text{M}^+ + \text{H}_2\text{O}
\]

\[
\text{acid} + \text{metal carbonate} \rightarrow \text{salt} + \text{H}_2\text{O} + \text{CO}_2
\]

\[
\text{HA} + \text{MCO}_3(s) \rightarrow \text{M}^+ + \text{H}_2\text{O} + \text{CO}_2(g)
\]

Hydrolysis is the reaction of a salt with water. During the hydrolysis, the salt will form an acidic, alkaline or neutral solution. The acidity or alkalinity of the salt is dependent on the relative strength of the acid and base that is used.

\[
\text{NaHCO}_3 \rightarrow \text{Na}^+ + \text{HCO}_3^-
\]

\[
\text{NH}_4\text{Cl} \rightarrow \text{NH}_4^+ + \text{Cl}^-
\]

\[
\text{Na}_2\text{CO}_3 \rightarrow 2\text{Na}^+ + \text{CO}_3^{2-}
\]

\[
\text{Na}_2\text{CO}_3 + \text{H}_2\text{O} \rightarrow \text{NaOH} + \text{H}^+ + \text{HCO}_3^-
\]

\[
\text{HCO}_3^- + \text{H}_2\text{O} \rightarrow \text{OH}^- + \text{H}_2\text{CO}_3
\]

\[
\text{H}_2\text{CO}_3 \rightarrow \text{H}_2\text{O} + \text{CO}_2
\]

\[
\text{H}_2\text{CO}_3 + \text{H}_2\text{O} \rightarrow 2\text{H}_2\text{O} + \text{CO}_2
\]

How to determine the pH of a salt

- \[
\text{NH}_4^+\text{Cl} \text{is a weak base}
\]

- Equilibrium lies to the left

- High concentration H^+

- \[
\text{NaHCO}_3 \text{is a weak acid}
\]

- Equilibrium lies to the right

- Low concentration H^+

- \[
\text{Na}_2\text{CO}_3 \text{is an alkaline}
\]

- Equilibrium lies to the right

- Low concentration OH^-

- \[
\text{NH}_4^+\text{Cl(aq)} \text{is acidic} \quad \text{pH}<7
\]

- \[
\text{NaHCO}_3\text{aq} \text{is alkaline} \quad \text{pH}>7
\]

TITRATIONS

A titration is a practical laboratory method to determine the concentration of an acid or base. The concentration of an acid or base can be determined by accurate neutralisation using a standard solution- a solution of known concentration. Neutralisation occurs at the equivalence point, when the molar amount of acid and base is the same according to the molar ratio.

\[
\frac{n_a}{n_b} = \frac{c_aV_a}{c_bV_b}
\]

\[n\text{- number of mole of substance (mol) / mol ratio} \]
\[c\text{- concentration of acid/base (mol.dm}^{-3}\text{)} \]
\[V\text{- volume of solution (dm}^3\text{)}

1 mL = 1 cm\(^3\)
1 L = 1 dm\(^3\)
1000 mL = 1 L
1000 cm\(^3\) = 1 dm\(^3\)

\[
\text{Hydrolysis is the reaction of a salt with water. During the hydrolysis, the salt will form an acidic, alkaline or neutral solution. The acidity of the salt is dependent on the relative strength of the acid and base that is used.}
\]

\[
\begin{align*}
\text{ACID} & \quad \text{BASE} & \quad \text{SALT} \\
\text{Strong} & \quad \text{Weak} & \quad \text{Acidic} \\
\text{Weak} & \quad \text{Strong} & \quad \text{Neutral} \\
\text{Weak} & \quad \text{Weak} & \quad \text{Alkali}
\end{align*}
\]

\[
\text{How to determine the pH of a salt}
\]

- \[
\text{NH}_4^+\text{Cl} \text{is a weak base}
\]

- Equilibrium lies to the left

- High concentration H^+

- \[
\text{NaHCO}_3 \text{is a weak acid}
\]

- Equilibrium lies to the right

- Low concentration H^+

- \[
\text{Na}_2\text{CO}_3 \text{is an alkaline}
\]

- Equilibrium lies to the right

- Low concentration OH^-

- \[
\text{NH}_4^+\text{Cl(aq)} \text{is acidic} \quad \text{pH}<7
\]

- \[
\text{NaHCO}_3\text{aq} \text{is alkaline} \quad \text{pH}>7
\]

INDICATORS

An indicator is a compound that changes colour according to the pH of the substance. During titrations, the indicator needs to be selected according to the acidity/alkalinity of the salt that will be produced (see hydrolysis).

\[
\begin{array}{|c|c|c|c|c|}
\hline
\text{INDICATOR} & \text{COLOUR IN ACID} & \text{COLOUR IN BASE} & \text{COLOUR AT EQUIVALENCE POINT} & \text{PH RANGE OF EQUIVALENCE} \\
\hline
\text{Litmus} & \text{Red} & \text{Blue} & \text{Pink} & 4,5 - 8,3 \\
\text{Methyl orange} & \text{Red} & \text{Yellow} & \text{Orange} & 3,1 - 4,4 \\
\text{Bromthymol blue} & \text{Yellow} & \text{Blue} & \text{Green} & 6,0 - 7,6 \\
\text{Phenolphthalein} & \text{Colourless} & \text{Pink} & \text{Pale Pink} & 8,3 - 10,0 \\
\hline
\end{array}
\]

Neutral: pH 7

Neutralisation is when the equivalence point is reached. Equivalence point is NOT when the solution is at pH 7, but when the molar amount of acid and base is the same according to the molar ratio. The pH at neutralisation is dependent on the salt that is formed. (see hydrolysis).

EXAMPLE:

During a titration, 25 cm\(^3\) of dilute H\(_2\)SO\(_4\) neutralises 40 cm\(^3\) of NaOH solution. If the concentration of the H\(_2\)SO\(_4\) solution is 0.25 mol.dm\(^{-3}\), calculate the concentration of the NaOH.

\[
\text{H}_2\text{SO}_4 + 2\text{NaOH} \rightarrow \text{Na}_2\text{SO}_4 + 2\text{H}_2\text{O}
\]

\[
n_a = 1 \quad c_a = 0.25 \text{ mol.dm}^{-3} \quad V_a = 25 \text{ cm}^3 = 0.025 \text{ dm}^3 \\

\]

\[
n_b = 2 \quad c_b = 7 \text{ mol.dm}^{-3} \quad V_b = 40 \text{ cm}^3 = 0.04 \text{ dm}^3 \\
\]

\[
\frac{n_a}{n_b} = \frac{c_aV_a}{c_bV_b} \\
\frac{1}{2} = \frac{(0.25)(0.025)}{(7)(0.04)} \\
\frac{c_b}{c_a} = 0.31 \text{ mol.dm}^{-3}
\]

By first principles:

\[
\frac{n_a}{n_b} = \frac{c_aV_a}{c_bV_b} \\
\frac{1}{2} = \frac{(0.25)(0.025)}{(7)(0.04)} \\
\frac{c_b}{c_a} = 0.31 \text{ mol.dm}^{-3}
\]

By titration equations:

Acid:

\[
\frac{n_a}{n_b} = \frac{c_aV_a}{c_bV_b} \\
\frac{1}{2} = \frac{(0.25)(0.025)}{(7)(0.04)} \\
\frac{c_b}{c_a} = 0.31 \text{ mol.dm}^{-3}
\]

Base:

\[
\frac{c}{V} = \frac{n}{V} = \frac{(0.25)(0.025)}{0.04} \\
\frac{c_b}{c_a} = 0.31 \text{ mol.dm}^{-3}
\]

For more information about Science Clinic’s seminars, classes and resources, visit www.scienceclinic.co.za
REDOX REACTIONS

Representing Redox Reactions

Redox reactions can be shown in two half-reactions showing the transfer of electrons.

Oxidation half-reaction: \( X \rightarrow X^{2+} + 2e^- \) (electrons are shown as products)
Reduction half-reaction: \( Y + 2e^- \rightarrow Y^2- \) (electrons are shown as reactants)

Net reaction: \( X + Y \rightarrow XY \) (no electrons are shown in the reaction)

Example:

In the following equations identify the oxidising and reducing agents:

\( \text{a) CuO + H}_2\text{O} \rightarrow \text{Cu} + \text{H}_2\text{O} \)
\( \text{b) } 2\text{NO} + 2\text{CO} \rightarrow \text{N}_2 + \text{CO}_2 \)

Step 1: Write all oxidation numbers
CuO: +2 H2O: +1 Cu: +2 H2O: +1

Step 2: Identify change in oxidation number for each element
Cu: +2 → 0 O: +2 → 0 H: +1 → +2

Step 3: Identify what has oxidised/reduced
\( \text{Cu}^{2+} (\text{CuO}) \): oxidation number decrease → gained electrons
\( \text{H}_2\text{O} \): oxidation number increase → lost electrons

Step 4: Write the half reactions
Ox: \( \text{CuO} + 2\text{e}^- \rightarrow \text{Cu}^{2+} \)
Red: \( 2\text{NO} + 4\text{e}^- \rightarrow 2\text{N}_2 \)

Step 5: Balance electrons
\( \text{Ox: } \text{CuO} + 2\text{e}^- \rightarrow \text{Cu}^{2+} \)
\( \text{Red: } 2\text{NO} + 4\text{e}^- \rightarrow 2\text{N}_2 \)

Step 6: Write overall reaction
\( 2\text{NO} + 2\text{CO} \rightarrow 2\text{N}_2 + \text{CO}_2 \)

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**LITHOSPHERE**

- The lithosphere is made up of the earth's crust and upper mantle.
- The crust contains non-renewable fossil fuels, minerals and soil chemicals (nutrients).

<table>
<thead>
<tr>
<th>ERA</th>
<th>PERIOD</th>
<th>MATERIALS AND WEAPONS</th>
<th>EVIDENCE FOUND IN SA</th>
</tr>
</thead>
<tbody>
<tr>
<td>Stone age</td>
<td>Early (2,5 mil - 200 000 years ago)</td>
<td>Hand axe - sharpened stone</td>
<td>Mrs Ples in Sterkfontein caves</td>
</tr>
<tr>
<td></td>
<td>Middle (250 000 - 30 000 years ago)</td>
<td>Stone treated with fire</td>
<td>Blombos caves</td>
</tr>
<tr>
<td></td>
<td>Late (30 000 years ago)</td>
<td>Tools made from stone, wood and bone. Pottery and copper objects</td>
<td>Melkhoutboom caves</td>
</tr>
<tr>
<td>Bronze age</td>
<td>30 000 - 1 800 years ago</td>
<td>Bronze tools and objects</td>
<td></td>
</tr>
<tr>
<td>Iron age</td>
<td>1 800 years ago</td>
<td>Iron metal weapons, more advanced</td>
<td>Mellville Koppies and Mapungubwe furnaces</td>
</tr>
</tbody>
</table>

- The elements in the lithosphere are not in element form. The compounds are in the form of ore- mixture of rock, sand and minerals.
- Mining and mineral extraction is required to extract useful materials from the lithosphere.

MINING IN SOUTH AFRICA

**GOLD MINING**

Gold is a precious metal that has a high market value.
Mined in the Witwatersrand (Gauteng), Welkom and Virginia (Free State), Klerksdorp (North West) and in Mpumalanga.

**STEPS IN GOLD EXTRACTION**

1. **Gold ore**
   Gold ore is extracted from deep-level mining.

2. **Comminution : Crushing and milling**
   Gold is crushed into a fine powder.

3. **Leaching: Cyanidation process**
   A cyanide solution is used to dissolve the gold from the fine powder.

   \[
   4\text{Au(s)} + 8\text{NaCN(aq)} + 2\text{H}_2\text{O(l)} + \text{O}_2(g) \rightarrow 4\text{NaAu(CN)}_2\text{(aq)} + 4\text{NaOH(aq)}
   \]

4. **Extraction**
   Gold is removed from solution by precipitation. Zinc is dissolved in gold solution, and zinc replaced the gold. The Gold forms a precipitate which is removed by filtration.

   \[
   \text{NaAu(CN)}_2\text{(aq)} + \text{Zn(s)} \rightarrow 2\text{Au(s)} + \text{Zn(CN)}_2\text{(aq)} + 2\text{NaCN(aq)}
   \]

5. **Recovery: Smelting**
   The precipitate that has been filtered out is smelted in a furnace to form gold bars of 99% purity.

6. **Refining: produces 99,9% pure gold**
   Calcination: other metals are oxidised in the presence of gold to remove moisture
   Fusion: Calcinated material is mixed with flux and melted in a furnace. Gold floats and is removed.

IMPACT OF GOLD MINING ON THE ENVIRONMENT

Gold mining results in holes in the ground. The soil removed creates mine dumps, resulting in a loss of biodiversity.
Water from mines can pollute ground and surface water.
Large amounts of energy is required to smelt the gold.
Cyanidation uses and forms cyanide compounds which are toxic.
Iron is a strong material used in manufacturing of a wide range of products. Large iron mines promote job creation and stimulates the economy. Iron is not mined in its pure form, it is found as magnetite ($\text{Fe}_3\text{O}_2$) and haematite ($\text{Fe}_2\text{O}_3$).

**Iron Mining**

Iron ore is mined from open-pit mines.

**Crushing**

Iron ore is crushed to reduce the particle size.

**Separation and Sorting**

Iron ore is separated according to purity. High grade iron ore are sent directly to blast furnace. Lesser grade iron ore undergos beneficiation.

**Beneficiation**

Lesser grade ore is further crushed and washed to form a slurry. The slurry is magnetised and pellets are formed from the slurry. Lastly the pellets undergo sintering, it is heated to set the pellets.

**Blending**

Pellets of various grades are mixed.

**Furnacing**

Iron ore pellets are placed into a blast furnace with coke (Carbon) and Limestone ($\text{CaCO}_3$). Iron ore is then reduced to pure molten iron. Impurities are removed as slag.

**Impact of Iron Mining on the Environment**

Open-pit mining destroys the landscape and vegetation.

Large amounts of water is required, reducing the amount of water in aquifers available to fauna and flora.

Large open-pit mines result in a lot of dust pollution in the atmosphere.

Large amounts of $\text{CO}_2$ are released from the blast furnacing processes.

Large amounts of electricity are required to melt the iron ore.

Iron products (steel etc.) are recyclable, however the recycling process required large amounts of electricity.

**Exploiting the Lithosphere**

**Iron Mining**

Iron is a strong material used in manufacturing of a wide range of products.

Large iron mines promote job creation and stimulates the economy.

Iron is not mined in its pure form, it is found as magnetite ($\text{Fe}_3\text{O}_2$) and haematite ($\text{Fe}_2\text{O}_3$).

Mined in Kathu (Northern Cape) and Thabazimbi (Limpopo).

**Iron Ore**

<table>
<thead>
<tr>
<th>Coke (C)</th>
<th>Limestone ($\text{CaCO}_3$)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Waste Gases</td>
<td></td>
</tr>
</tbody>
</table>

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Iron products (steel etc.) are recyclable, however the recycling process required large amounts of electricity.

**Phosphate Mining**

Phosphate is a primary nutrient for all plants, stimulating root development and increasing flower growth.

Phosphates are used to produce fertilisers, phosphoric acid and animal food supplements.

The largest phosphate mining site in South Africa is in Phalaborwa (Limpopo).

Phosphates are processed by Foskor (Richard's Bay) to produce phosphoric acid and phosphate-based fertilisers.

**Iron Ore**

<table>
<thead>
<tr>
<th>Coke (C)</th>
<th>Limestone ($\text{CaCO}_3$)</th>
</tr>
</thead>
</table>

**Impact of Phosphate Mining on the Environment**

Phosphate processing creates fluoride and $\text{SO}_2$ air pollutants.

The products of phosphates (fertilisers) can pollute water systems leading to eutrophication (algae bloom, algae die, decomposing algae removes oxygen out of the water, living organisms die).