SCH3UI
Making Observations and Taking Measurements

Observations

To notice with your **senses**. Senses may be aided by instruments such as rulers, microscopes, balances etc…

Inferences

To use **reason** and **knowledge** to make sense of your observations.

Example: The street is wet (**observation**). It rained last night (**inference**).

Observation - The fire alarm is going off.

Inference –

Observation – When a burning splint is placed in an unknown gas, the flame goes out.

Inference -

Types of Observations

Qualitative Observations:

Observations **describing** the nature of something **using your senses**. For example: colour, taste, texture etc…

DOES NOT INVOLVE NUMBERS!

Quantitative Observations:

Observations describing the **amounts** or **measurements** of something. For example: how fast, how hot, how much etc…

ALWAYS INVOLVE THE USE OF NUMBERS!

Describing matter

The properties that we can observe with our senses are called **physical properties.** The following is a list of some physical properties of matter that help us tell one thing from another.

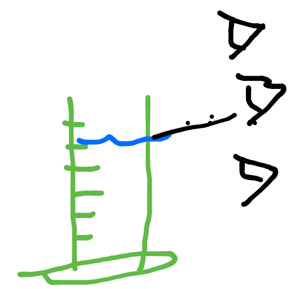
Physical Property	Explanation or Meaning
Physical State	solid, liquid or gas
Colour	black, white, colourless, greenish-blue, yellow
Odour	odourless, spicy, sharp, flowery
Taste	sweet, sour, salty, bitter
Clarity (transmission of light)	1. clear (transparent) 2. cloudy (translucent) 3. opaque (no transmission)
Lustre	ability to reflect light (shiny → dull)
Form (shape)	1. crystalline (regular shape, ex. salt) 2. amorphous (irregular shape, ex. pepper)
Texture	feel - fine, coarse, smooth, gritty
Hardness	scale [1 (soft, baby powder) → 10 (very hard, diamond)]
Brittleness	ability to shatter easily (not flexible)
Malleability	Can it be hammered into a sheet?
Ductility	Can it be stretched into a wire?
Viscosity	The resistance of a liquid to flowing. Syrup is viscous water is not.

Measurements

The study of chemistry often involves detailed measurements. Such quantitative observations always involve both a **quantity** and a **unit**. All measurements are limited by two main factors, the **equipment** and by the **skill of the person** using the equipment.

There is not much you can do to improve the piece of equipment that you are provided with, but you can determine how well you use that measurement tool!

How to Read a Measuring Tool Accurately



When taking measurements we should be aware of a few terms:

Parallax: The change in position of an object when the angle of view is changed.

Example: meniscus curve when reading a graduated cylinder.

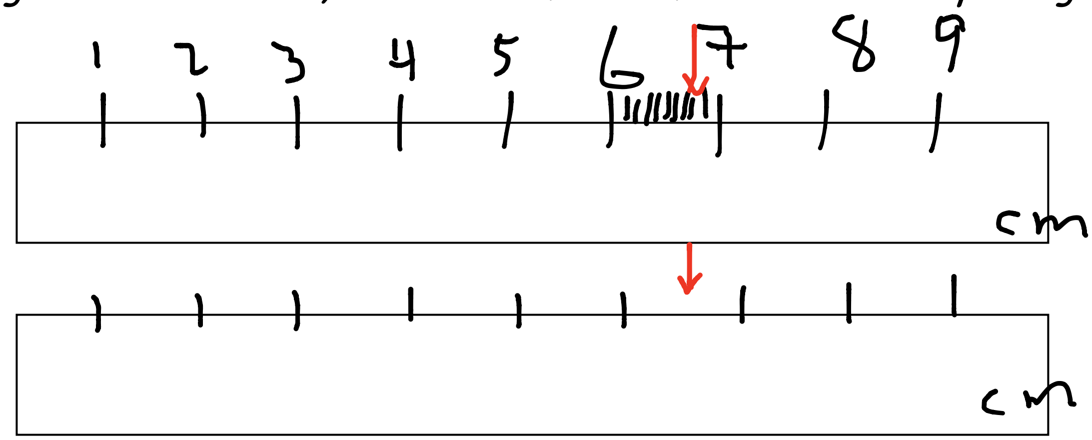
Accuracy: How close you are to a certain measurement.

Precision: The precision of a measurement is how sharply defined or detailed it is. A measurement can only be as precise as the scale used to measure it. The number of decimal places expresses precision.

Example: 1.50 cm is more precise than 1.5 cm.

If you record a measurement as 6.75 cm, you are saying that the 6.7 part of the measurement is absolutely certain. The third number "5" is an estimate. When measuring, you are allowed one estimation after the certain measurement. This is referred to as the level of uncertainty.

According to the ruler below, what is the measurement the arrow is pointing to?



6.82cm

How to Write Measurements

When recording a measurement, it is important to use proper format. Here is a list of guidelines to follow:

1. Every measurement consists of a number and a unit. A space is left between the last digit of a number and the unit. *Example: 75 cm not 75cm*
2. A period is not used after a unit abbreviation or symbol (except at the end of a sentence). *Example: 1.36 m not 1.36 m.*
3. Numbers and symbols must be used together. Numbers and words, or words and symbols are not used together. *Example: 6 m not 6 metres or six metres not six m*
4. Decimal fractions rather than common fractions are used. *Example: 0.25 cm not $\frac{1}{4}$ cm*
5. A zero is placed before a decimal marker. *Example 0.67 mm not .67 mm*
6. Long numbers are written with spaces separating groups of three digits (triads) on either side of the decimal marker. Numbers of four digits are not separated into triads.
   Example: 12 626.459 52 not 12626.45952
   2500 not 2 500

Converting Measurements

In the metric system, there are base units with which we make comparisons.

1. LENGTH - metres
2. MASS - grams
3. TIME - seconds
4. TEMPERATURE - °C
5. VOLUME - litres

In this system, prefixes are used to indicate the size of the base unit.

Use the conversion ladder to move between different units.

1. 5000 cm = 50.00 m
2. 0.005 kg = ____________ g
3. 8 mL = ____________ μL
4. 6 Ms = 600 000 das
5. 8.5 m = ____________ mm

Scientific Notation

For very large numbers and very small numbers scientific notation is used to avoid writing out many digits.

Example: 455 000 000 kg can be written as 4.55×10^8

Example: 0.000 26 m can be written as 2.6×10^{-4}

To convert standard notation to scientific notation:

1. Move the decimal so there is one non-zero digit in front of the decimal.
2. If the decimal was moved left the exponent is positive.
3. If the decimal was moved right the exponent is negative.
4. The number of movements is the exponent.

Write the following using scientific notation:

1. 580 000 ____________
2. 245 000 000 000 ____________
3. 23 000 000 000 000 ____________
4. 0.000 000 053 ____________
5. 0.0007 ____________
6. 0.000 0065 ____________

To convert from scientific notation to standard notation:

1. Move the decimal point to the right if the exponent is positive.
2. Move the decimal point to the left if the exponent is negative.

Write the following using standard notation:

1. 5.39×10^6 ____________
2. 9.8×10^4 ____________
3. 2.3×10^9 ____________
4. 2.25×10^{-5} ____________
5. 5.5×10^{-8} ____________
6. 9.3×10^{-12} ____________

Significant Digits

We can use significant digits as a simple method of providing estimates of the uncertainty in any measurement or calculation.

In the significant digit method, the uncertainty is presumed to be plus or minus 1 of the last digit reported. For example, if a volume is reported as 12.3 mL, there is an implied uncertainty of plus or minus 0.1 mL; the volume could be as small as 12.2 mL or as large as 12.4 mL. If the volume had been recorded as 12.30 mL, the uncertainty of the measurement would be plus or minus 0.01 mL.

To determine the number of significant digits reported in a measurement a few simple rules can be followed:

1. All non-zero digits (1, 2, 3 … 9) are significant 0.0001
2. Zeros that are located to the left of a value (leading zeroes) are **NOT** significant
3. Zeroes between other non-zero digits are significant 2002 (4)
4. Zeros that are located to the right of a value may or may not be significant (they can be used as placeholders)
5. The position of the decimal point is **NOT** important when counting significant digits

The following table shows the number of significant digits as well as the implied uncertainty for several measurements.

Measurement	Number of Significant Digits	Implied Uncertainty
307.0 cm	4	0.1
61 m/s	2	1
0.03 m	1	0.01
0.5060 km	4	0.0001
29.800 g	5	0.001
3.00×10^8 s	3	1 000 000

Using the rules for determining the number of significant digits and uncertainty, complete the following table.

Measurement	Number of Significant Digits	Implied Uncertainty
1.02 mm	3	0.01
203.45 mL	3	0.01
0.000 000 000 07 km	1	1×10^{-11}
2.8×10^5 g		
57.200 m/s		
386 L		
9.230 005 3 kg		
0.010 20 s		
7.28×10^9 m		

Any integer or exact number has an infinite number of significant figures. For example, the fact that there are 12 months in the year is an exact number. There is no level of uncertainty, and therefore an infinite number of significant digits.

The most challenging measurements to deal with when trying to determine significant figures are measurements such as 1500 cm. As previously mentioned, zeros to the right may or may not be significant. Therefore 1500 cm could have 2, 3, or 4 significant figures. It is impossible to tell without knowing how the measurement was obtained. To alleviate any ambiguity with such a measurement, scientific notation can be used.

2 significant digits = 1.5×10^3
3 significant digits = 1.50×10^3
4 significant digits = 1.500×10^3

Measurements and Calculations

When performing calculations using numbers that were obtained through measurement, there are two rules that must be followed to ensure that results do not show greater precision than the instruments used.

Certainty Rule

The certainty rule is used when **multiplying** and **dividing** measurement values. The number of **significant digits** in the answer **MUST** be equal to the **fewest** number of significant digits in any of the measurements.

Examples:

3 2 2 4 2

$1.20\ m \times 1.2\ m = 1.4\ m^2$

$5.603\ km \div 0.48\ h = 12\ km/h$

Precision Rule

The precision rule is used when **adding** and **subtracting** measurement values. The number of **decimal places** in the answer **MUST** be equal to the **fewest** number of decimal places in any measurement.

Examples:

1 3

520.2 mm + 88.221 mm = 608.4 mm

98.66 L – 64 L = 35 L

Perform the following Calculations:

9.55 cm + 2.1 cm = ________________

1248 km x 62.3 km = ________________

3 3

0.089 g – 0.008 g = 0.081 g

4 3

22.88 m ÷ 5.11 s = 4.48 s

Even/~~Odd Rounding~~

In the past, you probably have learned that when rounding numbers, when you encounter numbers 0-4 you round down and when you come across 5-9 you round up. Statistically speaking, this method is inaccurate. The problem is the number 5.

Imagine if you were charged 1% every time you made an Interac™ payment. One percent of $12.50, for example is $0.125 or 12 and a half cents. We don't have a half-cent piece, so should you be charged 13 cents? It is no big deal, right? Will it all even out in the long run? The answer is NO! So how do we avoid such errors? Even/odd rounding!!

Even/odd rounding is actually really easy. Let's just call it even rounding so you will remember what to do. Every time you are faced with a 5 as the first digit to be dropped and you are not sure if you should round it, just get rid of it and make sure the number to its left ends up being even.

For example, 4.75 rounded to two significant digits becomes 4.8 (see, the 8 is even!). The number 4.85 will also round to 4.8, so it again ends up an even number. This makes the process of rounding both consistent and fair!!

This "new" rule only applies when the number 5 is the last digit to be dropped; otherwise you still go about rounding as you have always been taught.

Examples:

Round the following numbers to 3 significant digits:

24.25 = 24.2

24.246 = ________________

24.35 = 24.4

24.658 = 24.7

24.750 = ________________

24.150 01 = ________________

You are now ready to begin making accurate observations and measurements!!!

Classifying Matter

Read pages 25-28 of your text and define the terms in the boxes below

Matter

Anything that has mass and takes up space (volume)

Pure Substance

A material that has definite composition

Mixture

A combination of two or more kinds of matter, in which each component retains its own characteristics

Element

A pure substance that cannot be broken down into smaller substances

$He_{(g)}$

Pb

Compound

A pure substance that is composed of two or more elements chemically combined in fixed proportions

H_2O

$NaCl$

homogeneous

Solution

A mixture in which the different components are mixed so they appear to be a single substance

salt water

heterogeneous

Mechanical Mixture

A mixture in which the different components can be distinctly seen

macaroni cheese

Physical and Chemical Changes in Matter

Physical Change – a change, such as change of state, which does not alter the composition of matter.

Chemical Change – the type of change that occurs when elements and/or compounds interact with each other to form different substances with different properties.

cooking eggs

Five Clues that a Chemical Change has occurred:

1. A new colour appears
2. Energy is given off (heat, light, sound etc…)
3. Bubbles of a new gas are formed
4. A solid material (precipitate) forms in a liquid
5. The change is difficult to reverse

Thought Lab: Mixtures, Pure Substances and Changes

Consider the following chemicals: table salt, water, baking soda, sugar, iron filings, sand, vegetable oil, milk and vinegar. Identify each chemical as a mixture or pure substance in the table below.

Pure Substance	Mixture

If you were to combine these different chemicals together, do you think you would produce a physical or a chemical change? Record your predictions in the table below.

	Table Salt	Water	Baking Soda	Sugar	Iron Filings	Sand	Vegetable Oil	Milk	Vinegar
Table Salt									
Water									
Baking Soda									
Sugar									
Iron Filings									
Sand									
Vegetable Oil									
Milk									
Vinegar									

P = Physical Change **C = Chemical Change**

An Introduction to the Periodic Table

During the mid 1800's, Russian scientist Dmitri Mendeleev invented the modern periodic table after noticing a relationship between the **physical** and **chemical properties** of the elements. He placed the elements in order of increasing **atomic mass.** At the time approximately **58** elements had been identified.

The modern periodic table, which is comprised of over **110** elements, **92** of which are naturally occurring, is organized by **atomic number** and makes use of element **symbols** that are the same throughout the entire world.

Metals are located on the **left-hand side** and throughout the **middle** of the Periodic Table. Metals are one kind of element that have certain **properties** in common –**malleable**, **ductile**, have **lustre**, **good** conductors of **heat** and **electricity**. All metals are **solid** except for **mercury** (Hg), which is a **liquid**.

Non-metals are located on the **right hand side** of the Periodic Table. Non-metals are **brittle**, not **ductile**, not very **shiny** and **poor** conductors of **heat** and **electricity**. At room temperature non-metals may be **solids** or **gases** and one, **bromine** (Br), is a liquid.

A division line known as the "**staircase**" separates metals and non-metals. On either side of the staircase are a group of elements known as **metalloids** that show **characteristics** of both metals and non-metals.

The name for each **horizontal** row of the Periodic Table is a **period**. There are **seven** periods.

The **vertical** columns in the periodic table are called **groups** and range from 1-18 (these are typically written as Roman Numerals). Some groups are given special names because they form a **family** of elements with **strong relationships.**

There are four families within the periodic table:

Group 1 – **Alkali Metals**
Group 2 – **Alkaline Earth Metals**
Group 17 – **Halogens**
Group 18 – **Noble Gases**

Atoms and Their Composition

Elements are the basic substances that make up all **matter**.

An atom is the smallest particle of an element that still retains the **identity** and **properties** of the element.

Atoms are made up of even smaller particles. These **subatomic** particles are ***protons***, ***neutrons*** and ***electrons***.

Protons and neutrons make-up the **nucleus** or core of an atom and contribute to the **mass** of an atom, while electrons are **fast moving** and occupy the **space** that surround the nucleus of the atom (**orbitals**). Electrons are so **small** and **light** that they essentially contribute no overall weight to the atom.

Subatomic Particle	Charge	Symbol	Mass (g)	Radius (m)
Electron	1-	e^-	9.02×10^{-28}	Smaller than 10^{-18}
Proton	1+	p^+	1.67×10^{-24}	10^{-15}
Neutron	0	n^0	1.67×10^{-24}	10^{-15}

Since subatomic particles are so light, chemists use a unit called an **atomic mass unit (u)** for their measurement. Both protons and neutrons have a mass of **1 u.**

Every Element has a unique:

Mass
Number **A**
<u>X atomic symbol</u>
Atomic
Number **Z**

- **Name**
- **Symbol**
  Capital letter, followed by one or two lowercase letters if present; each symbol is unique
- **Atomic number (Z)**
  Equals the number of protons in the nucleus. It is inferred that the number of electrons is the same since an element is electrically neutral
- **Atomic Mass (A)** big # rounded
  Equals the number of protons and neutrons in the nucleus

We can use this information to calculate the number of neutrons by means of the following equation:

Number of neutrons = Mass number (A) – Atomic number (Z)

Examples:

$^{39}_{19}K$ $\quad \# n^\circ = 39 - 19$

$= 20 n^\circ$

You will notice that an element reports an **atomic weight** (a decimal number) instead of a mass number on the Periodic Table. The atomic weight represents a "**weighted average**" of all the **isotopes** for a particular atom.

Isotopes are atoms of an element that have the same number of **protons** in their nucleus, but a different number of **neutrons**.

Isotopes have very similar **chemical** properties, but they differ in **physical** properties.

Example:

"Light" Lithium	"Heavy" Lithium
$^{6}_{3}Li$	$^{7}_{3}Li$
7.42%	92.58%

avg atomic mass $= (\text{amt Li-6})(\text{mass Li-6}) + (\text{amt Li-7})(\text{mass Li-7})$

The **mass number** of an atom can be determined by **rounding** the **atomic weight** on the Periodic Table.

$= (0.0742)(6) + (0.9258)(7)$

$= 6.93$ amu

Examples:

17 Cl 35.45	7 N 14.01	18 Ar 39.95

N^{-3}

p^+	n°	e^-
17	18	17

$7p^+ \quad 7n^\circ \quad 10e^-$

How to Draw Atoms

Draw Bohr-Rutherford Diagrams

Ernest Rutherford and Niels Bohr developed the planetary model of the **atom** in 1913. In this model, the nucleus, containing the **protons** and **neutrons**, takes the central place just like the Sun takes the central place in our solar system. The electrons spin around the nucleus in orbits similar to the path of the planets around the Sun. The orbits represent the different amounts of **energy** that the **electrons** can have. Electrons in the first orbit have the **least** energy, whereas electrons in the last orbital have the **most** energy. The first orbit holds up to **2** electrons. The second and third orbits contain up to **8** electrons. As you fill the orbits, always fill the **lowest** energy orbit first, then fill up the next one and the next and so on.

When you draw Bohr-Rutherford diagrams of an element, you identify the **number** of **protons** and **neutrons** in the center of the atom and place **dots** to represent the **electrons** in their orbits. Since electrons have a **negative** charge, and according the law of **electrostatics**, **oppositely** charged particles **attract** and **like** charges **repel**; you must place the first **four** electrons in the orbit as far apart as possible. For reasons beyond the scope of this course, the next **four** electrons in the orbit (if there are any) pair up with the electrons already there.

31
15 P

Step 1: Determine the number of protons

This is equal to the atomic number of the element

$15p^+$

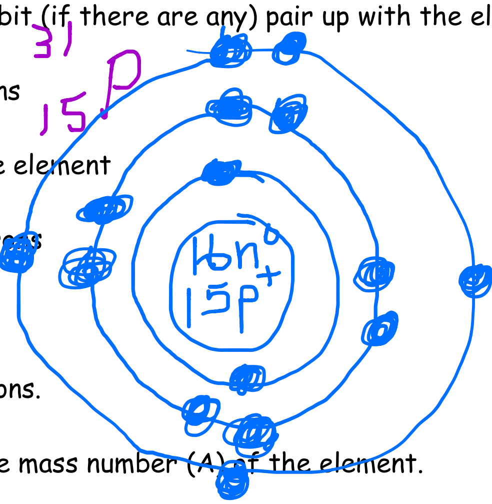
Step 2: Determine the number of electrons

$15e^-$

This is equal to the number of protons.

Step 3: Determine the number of neutrons.

$31 - 15 = 16n^0$



Subtract the atomic number (Z) from the mass number (A) of the element.

$16n^0$ $15p^+$ — $2e^-$ — $8e^-$ — $5e^-$

Step 4: Draw a nucleus and write in the number of protons and neutrons.

Step 5: Draw electron shells around the nucleus and fill them with the appropriate number of electrons. Always fill the inner shells to their maximum before moving to the outer shells.

<u>Lewis (Electron) Dot Diagrams</u>

Lewis Dot Diagrams are a short way to show the **last** energy shell (**valence**) shell for an atom. These are the electrons on the outer perimeter of an atom and generally the ones that will be involved in **bonding**.

The element **symbol** is used to represent, the **proton**, **neutrons** and all **inner electrons**. Just like when drawing B/R Diagrams, the first four valence electrons (dots) should be drawn as far apart as possible, one on each side of the element **symbol**. The remaining four electrons (if present) can then be paired up.

Patterns and Trends in the Periodic Table

Valence Electrons

You can determine the number of **valence electrons** in any of the **Main Group Elements**, without fully drawing a Bohr-Rutherford Diagram by looking at the **group** number that the element is in.

Group Number	1	2	13	14	15	16	17	18
Number of Valence Electrons	1	2	3	4	5	6	7	8

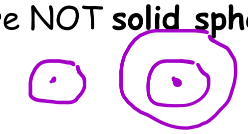
Energy Levels

You can determine the number of **energy levels** (orbitals) an atom contains by looking at the **period** number that the element is in.

Period Number	1	2	3	4	5	6	7
Number of Energy Levels	1	2	3	4	5	6	7

Atomic Size or Radius

Chemists measure an atom's size in terms of its **radius** in **picometres** (pm), which is **10^{-12} m**. The atomic radius is the distance from the **nucleus** to the approximate outer boundary of **electrons**. The outer boundary is only an approximation because atoms are NOT **solid spheres**, but more closely resembles a **cloud-like** structure.



As you move **down** a group the atomic radius **increases**. Each additional period represents an extra **energy level**, therefore the atom is continuously becoming bigger as the valence electrons occupy a level that is farther from the nucleus. The total number of **shells** that an atom has is equal to the **period** number where it is positioned within the Periodic Table.

As you move **across** a period the atomic radius **decreases**. Elements in the same period have the same number of energy levels, however, the **positive** charge in the nucleus **increases** across the period. This positive force **pulls** the outer electrons closer, reducing the atom's total size.

→ atomic #

Zeff

} effect of the nuclear CHARGE

↳ yo, how many p+ you got?

Complete Practice Problem #7 on Pg. 52

a) K, Ca, Sr Ca, K, Sr	f)
b)	g)
c) Rb, B, I B, I, Rb	h)
d)	i)
e)	j)

Ionization Energy

The particle that results when a neutral atom gains or gives up electrons is called an **ion**. There are two types of ions, a **cation** (positive) and an **anion** (negative). Atoms tend to form ions in an attempt to achieve a **stable octet**.

Metals in the main group elements tend to **give up** electrons forming **cations** that have the same number of electrons as the nearest **noble gas** (located in the previous period).

Non-metals tend to **gain** electrons and form **anions** that have the same number of electrons as the nearest **noble gas** (located within the same period).

Ionization energy is the energy required to **remove** an electron from an atom. The energy required to remove the first electron from a stable atom is called the **first ionization energy**. It is measured in **KJ/mol**, where **KJ** is a unit of energy and **mol** is an amount of a substance (you will learn more about the mol in chapter 5).

Ionization energy tends to **decrease down a group**. As you move down a group the number of energy levels **increase**. Valence electrons therefore become **farther** away from the positive attractive force of the nucleus, and are **easier** to remove.

Ionization energy tends to **increase across a period**. The attraction between the nucleus and valence electrons increases as more **protons** (positive charges) are added to the nucleus. Thus, more **energy** is needed to pull an electron away from its atom.

Cs Mg S   Cs Mg S

Complete Practice Problems #8-9 on Pg. 55

a)	d)
b)	e)
c)	f)

a)	d)
b)	e)
c)	f)

Electron Affinity

Electron affinity is a measure of the **change in energy** that occurs when an electron is **added** to the outer energy level of an atom.

Negative electron affinity → energy is **released** when an atom gains an electron

Positive electron affinity → energy is **absorbed** when an atoms gains an electron

Electron affinity tends to **decrease down a group**. As you move down a group the number of energy levels increase. With a greater distance from the nucleus, incoming electrons are NOT readily attracted.

Electron affinity tends to **increase across a period**. The attraction between the nucleus and electrons increases as you move across a period. With a stronger attractive force, incoming electrons are more readily drawn towards the atom.

The units for electron affinity are the same as the units for ionization energy, **KJ/mol**.

Complete Section Review #2-4 on Pg. 60 in the space below

Classifying Chemical Compounds

Recall: a compound is a pure substance composed of two or more elements, chemically bonded in fixed proportions.

Chemical bonds are forces that attract atoms to each other. **Bonding** involves the interaction between the **valence electrons** of atoms and is the driving force of **stability**.

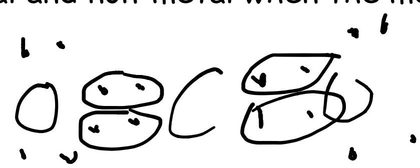
While there are only **90** naturally occurring elements, there are **thousands** of different compounds. To help organize these compounds, chemists classify them into two main groups based on the **type** of bond that they form, and according to their **properties**.

Ionic Bond

A chemical bond between **oppositely** charged **ions** that arise from the **transfer** of **electrons**. It usually involves a **metal** (low ionization energy) and a **non-metal** (**high electron affinity**).
Example: **NaCl** $[Na]^{+1}$ $[:Cl:]^{-1}$

Covalent Bond

A chemical bond in which **electrons** are **shared** by two atoms. It usually involves two **non-metals** (Example: CO_2), but can also occur between a metal and non-metal when the metal has a fairly high **ionization energy**.

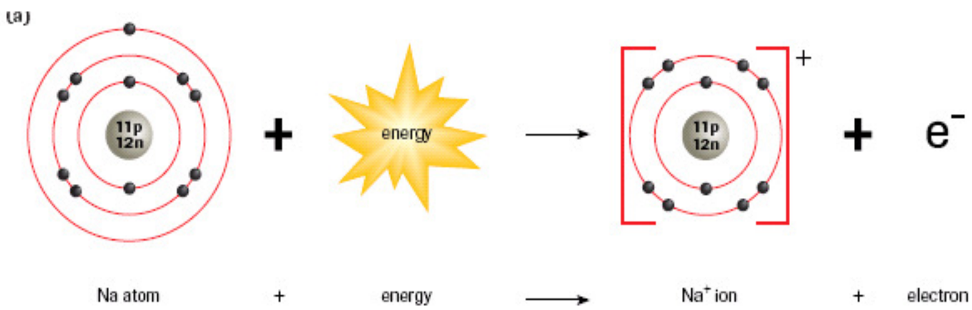


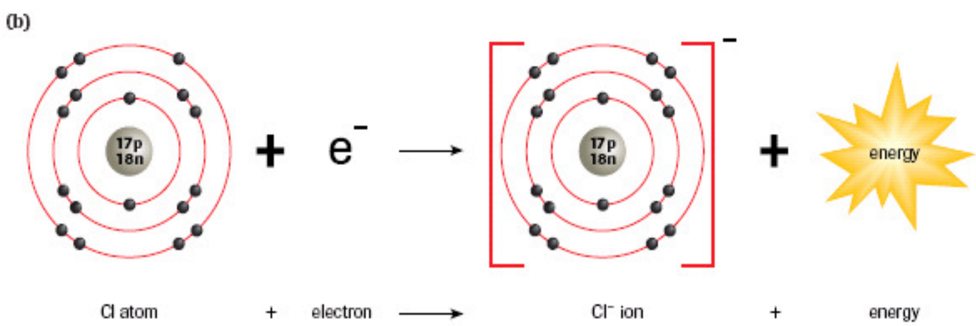
Comparing Ionic and Covalent Compounds

Property	Ionic Compound	Covalent Compound
State at room temperature	Crystalline solid	Liquid, gas, solid
Melting point	High	Low
Electrical conductivity as a liquid	Yes	No
Solubility in water	Most have high solubility	Most have low solubility
Conducts electricity when dissolved in water	Yes	Not usually

Ionic Compounds

In order for an ionic compound to form, an atom must first become an **ion**. To do this an atom will either **gain an electron(s)** to become an **anion** or **loose an electron(s)** to become a **cation**. An anion has a **negative** charge and a cation has a **positive** charge. Anions and cations will **attract** to one another, forming an **ionic bond**. Atoms will **exchange** their electrons and form ions in order to get a **full valence** shell of electrons (**stable octet**).





When two atoms or ions have the same electron configuration they are said to be **isoelectronic**.

Example: O^{-2} is isoelectric with **Ne**. Both have **10 e⁻** and a full energy level.

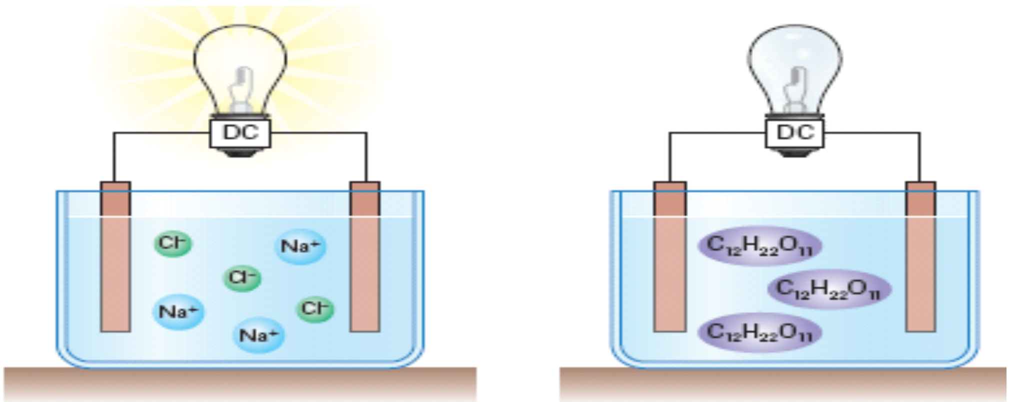
What happens when lithium becomes an ion? What is it isoelectric with?

$Li + energy \rightarrow [Li]^{+1} + e^-$ He

What happens when nitrogen becomes an ion? What is it isoelectric with?

$N + 3e^- \rightarrow [N]^{-3} + energy$ Ne

Ionic compounds are capable of **conducting electricity**. This is because they are composed of **ions**. Electricity is the movement of **negatively charged** particles. Ionic solids are **NOT** able to conduct electricity because the ions are held in place in a **rigid crystal lattice configuration,** which prevents the ions from moving to any degree. When **melted** or **dissolved** in water, the ions will split apart from each other (**dissociate**) and are then free to move around. A substance that can conduct electricity is termed an **electrolyte**.



You cannot always identify if a compound is ionic or covalent via its **properties**. A second method that can be employed to determine the type of bond in a compound is to examine the bonding atom's **electronegativity**.

Electronegativity (EN) is a measure of an atom's ability to attract electrons **in a chemical bond. There is a specific electronegativity associated with each element.**

The differences between electronegativities (**ΔEN**) can be used to determine if a bond is ionic or covalent. When calculating ΔEN, the **smaller** value is always subtracted from the **larger** value, so the answer is always **positive**.

Chemists consider bonds with a ΔEN greater than **1.7** to be **ionic** and bonds with a ΔEN less than **1.7** to be **covalent**.

If the ΔEN dictates that atoms will form an ionic bond, we can use electron dot diagrams to show the movement of electrons and the resulting ions that will form.

Bonding Atoms	ΔEN	Type of Bond	Formation of Bond (Movement of Electrons)	Ions formed	Chemical Formula
Lithium and Bromine	Br 2.96 Li 0.98 1.98	ionic	$Li + Br$	$[Li]^{+1}$ $[Br]^{-1}$	$LiBr$
Magnesium and Oxygen					
Beryllium and fluorine					
Aluminium and Sulphur	S 2.58 Al 1.61 0.97	ionic (weak)	$Al + S$ $+ S$ $Al \quad S$	$2[Al]^{+3}$ $3[S]^{-2}$	Al_2S_3

Covalent Compounds

Covalent compounds typically form when two or more **non-metals** bond together. During a covalent bond, valence electrons are **NOT** exchanged, but rather are **shared** between atoms. Atoms can share **one pair** of electrons, creating a **single** bond; **two pairs** of electrons, creating a **double** bond; or **three pairs** of electrons resulting is a **triple** bond. Atoms will share as many electrons as they need in order to achieve a stable octet.

Multiple Covalent Bonds

One pair of electrons shared	→	Single bond	→	$H\cdot + \cdot H \rightarrow H:H$
Two pairs of electrons shared	→	Double bond	→	$O + O \rightarrow O::O$
Three pairs of electrons shared	→	Triple bond	→	$N + N \rightarrow N:::N$

partial charge – δ

$+\!\!\longrightarrow$ dipole moment where e⁻ @

Similar to what we discovered about ionic compounds, we can determine if a covalent bond will form by looking at an atom's **electronegativity**. Recall, that if the ΔEN is **greater** than **1.7**, there will be an **exchange** of electrons, **ions** will form and an **ionic compound** will be produced. A covalent compound will occur if the ΔEN is **less** that **1.7**. But there are actually, three different types of covalent bonds – **pure**, **non-polar** and **polar**. The ΔEN's listed below outline when each type of covalent bond will occur.

ΔEN: pure covalent | non-polar covalent | polar covalent | ionic

0 δ^+ δ^- 0.4 δ^+ δ^- 1.7 $[\ ]^+ [\ ]^-$ 3.3

When a pair of electrons is shared between atoms of the **same element**, the ΔEN will always be **0** and they form what is called a **pure covalent bond**. In this arrangement, the electrons are shared **exactly in-between** the two atoms. There are only seven such elements that occur naturally; they are called **diatomic** molecules: H_2, O_2, F_2, Br_2, I_2, N_2, Cl_2

When two different atoms have a **similar attraction** for electrons, (ΔEN between 0.1 - 0.5) then the atoms will also share their electrons **relatively equally** and result in the formation of a **non-polar covalent bond**.

$\overset{\delta^-}{C} \longleftarrow \overset{\delta^+}{H}$

ΔEN
C 2.55
H 2.20

0.35

The sharing of electrons between atoms will **NOT** always be **equal**. When one atom in a bonded pair has a **stronger attraction** for the electrons (ΔEN between 0.5 – 1.7), the shared pair of electrons may spend more **time** around one atoms than around the other. When an electron pair is **NOT** shared equally, there is a **partial** negative charge around one atom and a partial positive charge around the other; this is referred to as a **partial dipole** and is represented by the symbol δ. The bond between the two atoms is called a **polar covalent bond**.

Predict the type of bond that will occur between the following:

	Atoms	ΔEN	Bond	Diagram (show bond polarity if required)
1	Carbon and Bromine	Br 2.96 C 2.55 0.41	$\delta^+ \rightarrow \delta^-$ C — Br	polar covalent
2	Nitrogen and fluorine			
3	Phosphorus and sulfur			
4	Oxygen and Iodine	O 3.44 I 2.66 0.78	polar covalent	$\delta^- \leftarrow \delta^+$ O — I

Low Conductivity of Covalent Compounds

Recall: ions are required for conductivity

Covalent compounds form **solids**, **liquids** or **gases**. The atoms in each compound are held together by strong covalent bonds and thus do not **break-up into ions** when they **melt** or **boil**, but rather remain as **intact molecules**. For this reason, covalent compounds are also called **molecular compounds**.

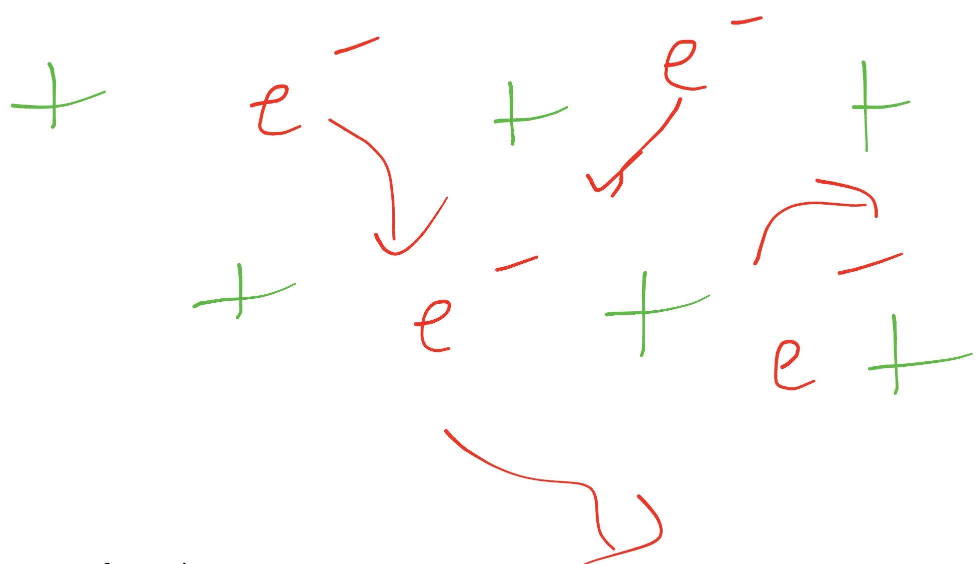
Metallic Bonding

How do metals bond?

- They do not form **ionic bonds** (ΔEN is not greater than 1.7)

- The do not form **covalent bonds** (there are not enough valence electrons to be shared to form a stable octet)

In metallic bonding, atoms release **electrons** to a **shared pool** of electrons. Essentially the **metal ion** is in a sea of **electrons**, which does not take on a particular **orientation**.



This explains the many properties of metals:

Conductivity $\rightarrow$ many e^- free to move everywhere

Malleable $\rightarrow$ no particular shape, e^- can slide past one another

Electronegativity

You cannot always identify if a compound is ionic or covalent via its **properties**. A second method that can be employed to determine the type of bond in a compound is to examine the bonding atom's **electronegativity**.

Electronegativity (EN) is a measure of an atom's ability to attract electrons in a chemical bond. There is a specific electronegativity associated with each element (figure 3.6, Pg. 71).

Similar to atomic radius, ionization energy and electron affinity, electronegativity shows definite **trends** in the Periodic Table. Radius, ionization and affinity refer to properties of **single atoms**, while electronegativity refers to properties of atoms that are involved in **chemical bonding**.

*Electronegativity tends to **decrease down a group**. Increasing energy levels between valence electrons and the nucleus means that there is **less attraction** between the nucleus and bonding electron pairs.*

*Electronegativity tends to **increase across a period**. Due to a smaller atomic radius, atoms attract a bonding pair of electrons with **greater strength**, because the bonding pair can move **closer** to the nucleus.*

The differences between electronegativities (**ΔEN**) can be used to determine if a bond is ionic or covalent. When calculating ΔEN, the **smaller** value is always subtracted from the **larger** value, so the answer is always **positive**.

Chemists consider bonds with a ΔEN greater than **1.7** to be **ionic** and bonds with a ΔEN less than **1.7** to be **covalent**. Covalent bonds that have a ΔEN between **1.7** and **0.~~5~~4** are termed **polar covalent**, while those that are between **0.~~5~~4** and **0.1** are **non-polar covalent**. If a covalent bond has a ΔEN of **exactly zero**, it is termed a **pure covalent bond**.

pure, non-polar, polar, ionic

0 0.4 1.7 3.3

TEST(4)

Examples:
Determine the ΔEN for each bond shown. Indicate whether each bond is ionic or covalent.

a) NO_2 b) CaS d) I_2 e) CCl_4

O 3.44
N 3.04
0.40

polar covalent

$\delta^+ \longrightarrow \delta^-$

N—O

Ionic and Covalent Bonding: The Octet Rule

Octet Rule - atoms bond in order to achieve an electron configuration that is the same as the electron configuration of a noble gas.

Recall that noble gases are the most **stable** elements in the Periodic Table. Their **extremely unreactive** nature can be attributed to the fact that they have a **full** outer electron level.

When two atoms or ions have the same electron configuration they are said to be **isoelectronic**.

Example: O^{2-} is isoelectric with **Ne**. Both have **10 e^-** and a full energy level

Stable Octets and Ionic Bonds

When the ΔEN between two atoms is **greater** than 1.7, the resulting bond that forms is **ionic**. In attempt to gain a stable octet, these atoms **transfer** electrons.

What happens when lithium and bromine bond?

The bond is ________. Lithium becomes isoelectric with ____ and bromine with ____.

What happens when magnesium and oxygen bond?

The bond is ________Magnesium becomes isoelectric with ____ and oxygen with ____.

Practice Problem #2, Pg. 76

Bonded Atoms	ΔEN	Type of Bond	Bonded Atoms	ΔEN	Type of Bond
a)			d)		
b)			e)		
c)			f)		

Practice Problem #3, Pg. 76

a)	d)
b)	e)
c)	f)

What happens when beryllium and fluorine bond?

The bond is ________. Beryllium becomes isoelectric with _____ and fluorine with _____.

Practice Problems #4 Pg. 78

Bonding Pair	ΔEN	Bonding Pair	ΔEN
a)		d)	
b)		e)	
c)		f)	

Practice Problems #5 Pg. 78

a)	d)
b)	e)
c)	f)

Conductivity of Ionic Compounds

An electrical current can flow only if charged particles are available to **move** and **carry** the **current**.

Ionic compounds in a solid state **CANNOT** conduct. This is because the ions are arranged in a **rigid lattice formation**, which prevents the ions from moving to any degree.

Ionic compounds in a **liquid** state, or that have been **dissolved** in water **CAN** conduct. This is because the ions are **spread further apart** and free to move.

Stable Octets and Covalent Bonds

When the ΔEN between two atoms is **less** than 1.7, the resulting bond that forms is **covalent**. In attempt to achieve a stable octet, these atoms **share** electrons because neither atom has a noticeably **higher affinity** for electrons. In other words, the electrons are equally attracted to each atom.

When an electron is shared between atoms of the same element **(ΔEN = 0)**, they form what is called a **pure covalent bond**. There are seven such elements that occur naturally, and they are called **diatomic** molecules: **H_2, O_2, F_2, Br_2, I_2, N_2, Cl_2**

Covalent bonds can be shown using Lewis Structures:

Other elements also have covalent bonds even though ΔEN is slightly greater than 0:

Practice Problem #6, Pg. 81

a)	c)
b)	d)

Practice Problem #7, Pg. 81

a)	b)	c)
d)	e)	f)

Multiple Covalent Bonds

Just like some atoms transfer more than one electron in an ionic bond, sometimes atoms need to share two or three pairs of electrons to achieve a stable octet.

One pair of electrons shared	→	Single bond	→
Two pairs of electrons shared	→	Double bond	→
Three pairs of electrons shared	→	Triple bond	→

Practice Problem #8-10, Pg. 82

8.
9.
10.

Low Conductivity of Covalent Compounds

Recall: ions are required for conductivity

Covalent compounds form **solids**, **liquids** or **gases**. The atoms in each compound are held together by strong covalent bonds and thus do not **break-up into ions** when they **melt** or **boil**, but rather remain as **intact molecules**. For this reason, covalent compounds are also called **molecular compounds**.

Polar Covalent Interactions

Polar Bonds

How can we explain the wide variety of properties that covalent compounds have (solids, liquids, gases etc…)?

For example:

*Both CO_2 and H_2O are covalent compounds, yet carbon dioxide is a **gas** at room temperature and water is a **liquid**. Why?*

*H_2O has a boiling point of **100°C**, while N_2O has a boiling point of **-89°C**. Why?*

The differences in the properties of these compounds are explained by the **ΔEN** of their bonds.

When two bonding atoms have a ΔEN that is greater than 0.5 but less than 1.7 they form what is called a **polar covalent bond**.

This ΔEN is not large enough to generate a **transfer** of electrons, but does cause the electron pair to spend **more time** near the more **electronegative** atom.

Practice Problems #11-13 Pg. 86

a.	c.	e.	g.
b.	d.	f.	h.

a.	c.	e.	g.
b.	d.	f.	h.

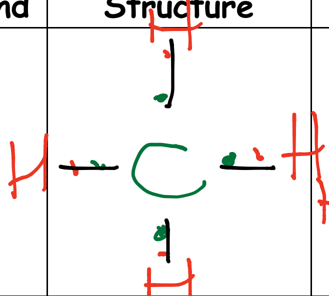
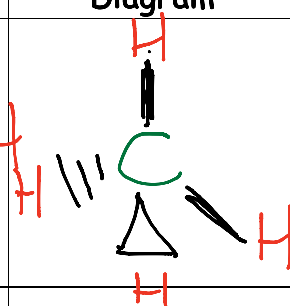
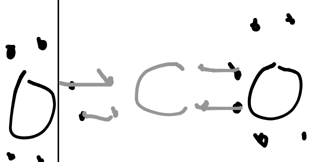
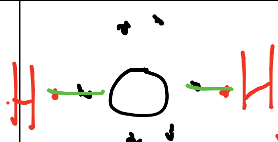
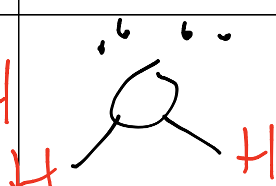
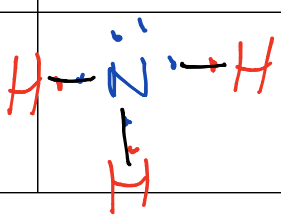
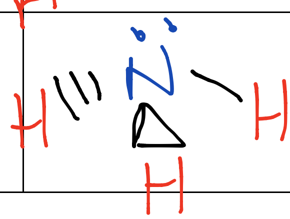
a.
b.

Molecular Shape

Not only can the covalent bonds between atoms be considered polar or non-polar, but also a **molecule** itself can show these characteristics. The polarity of a molecule is dependent upon the **shape** of the molecule. Molecular compounds take on a variety of shapes depending upon the arrangement of their **bonding pairs** (e^- pairs involved in a bond) and **lone pairs** (e^- not involved in bonding). Lone pairs especially play a large role in the shape of a molecule.

All pairs of electrons find an arrangement so that they are as **far away** from each other as possible. Remember, electrons are negatively charged and will **repel** each other.

There are a variety of diagrams to represent the arrangement and shape of molecules such as **Lewis Structures**, **Structural Diagrams**, **Ball-and-Stick Models** and **Space-Filling Models**. (See page 87, Figure 3.26 for examples)

Covalent Compound	**Lewis Structure**	**3-D Structural Diagram**	**Shape**	**Explanation**
CH_4			Tetrahedral 4 BP 0 LP	This shape allows for bonding pairs to be the maximum distance apart
CO_2			Linear 2 BP 0 LP 1 BP	No lone pairs, so the two bonding pairs arrange themselves opposite to one another
H_2O			Bent 2 BP 2 LP	Four electron pairs, two of which are lone
NH_3			Pyramidal 3 BP 1 LP	Four electron pairs, one of which is lone

Polar Molecules

To determine if a molecule is polar, you must first examine if the **bonds** themselves are polar by comparing the **electronegativites** of bonding atoms and then look to the **shape** to molecule.

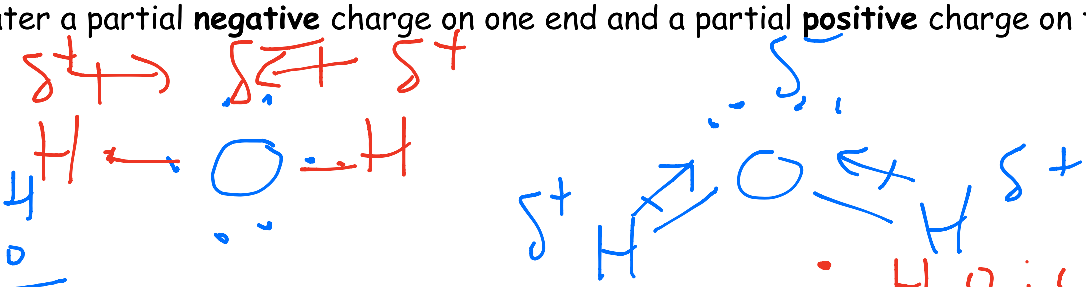
1. @least one polar covalent bond
2. Is the molecule asymmetrical

For example:

Water has polar bonds between its atoms and is considered a **polar** molecule, because its molecular shape is **bent**. This **asymmetrical** structure (caused by the **lone pair** of electrons) gives water a partial **negative** charge on one end and a partial **positive** charge on the other.

H_2O

ΔEN
O 3.44
H 2.20
1.24



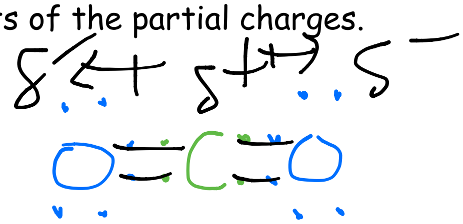
1. ✓
2. ✓

∴ H_2O is a polar molecule

Carbon dioxide is a **non-polar** molecule, because the molecule is **linear**. Even though there are polar bonds between the carbon and the oxygen atoms, the **symmetrical** arrangement cancels out the effects of the partial charges.

CO_2

ΔEN
O 3.44
C 2.55
0.89



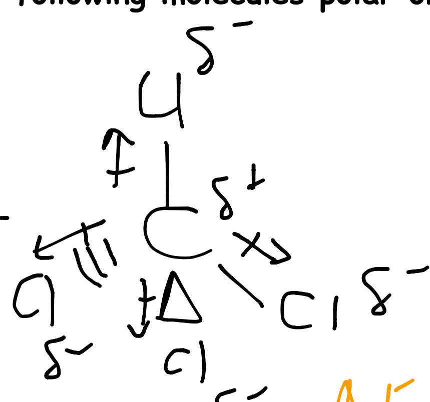
1. ✓
2. ✗

∴ CO_2 is a non-polar molecule

Are the following molecules polar or non-polar?

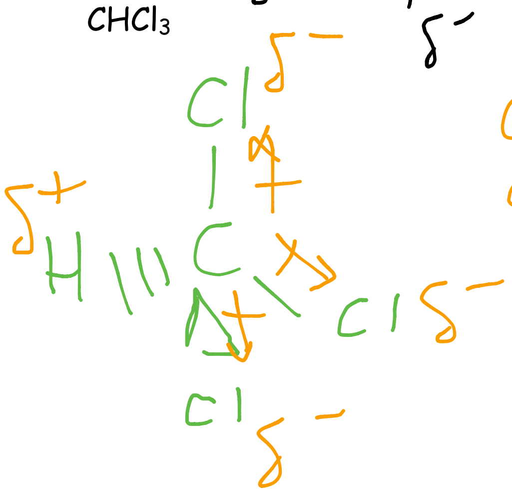
CCl_4

Cl 3.16
C 2.55
0.61



1. ✓
2. ✗ non-polar molecule

$CHCl_3$



ΔEN
Cl 3.16
C 2.55
0.61

C 2.55
H 2.20
0.35

1. ✓
2. ✓

∴ $CHCl_3$ is a polar molecule

bonds

Intermolecular and Intramolecular Forces

Going back to our original question of how can we explain the wide variety of properties that covalent compounds have (solids, liquids, gases etc)?

The answer is polarity!!!

Recall that ionic compounds are held together through **electrostatic** interactions, which create **solid crystal lattice** structures. Covalent compounds can be solids, liquids or gases at room temperature, which means there must be something **different** that holds these molecules together and creates such variation. These forces are called **intermolecular** forces and occur in molecules that show **polarity**.

Intermolecular Forces (physical bonds):
The forces that exist **between** molecules (ie. **Hydrogen** bond).

Intramolecular Forces (Chemical bonds):
The forces that bond **atoms to each other** within a molecule (ie. **Covalent** bond).

There are several different types of intermolecular forces, for now we will just examine one; the hydrogen bond, which can explain all of the properties that water exhibits, such as why water is a **liquid** at room temperature, why it "**sticks**" to itself and can account for its **high** melting and boiling points.

Since water is a polar molecule, it has one end, which has a partial positive charge, and the other end, which has a partial negative charge. When two water molecules come into close contact, there will be **weak attractive forces** between the partially negative **oxygen** of one molecule and the partial positive **hydrogen** of the other. This intermolecular attraction is called a hydrogen bond. A hydrogen bond can occur in any polar molecule where hydrogen is present.

Non-polar molecules do not attract each other as much as polar ones and consequently have **lower** melting and boiling points, and are typically **gases** at room temperature.

bonds

Intramolecular Forces and Intermolecular Forces

Intramolecular forces are the forces of attraction within molecules. Examples: non-polar covalent bonds, polar covalent bonds, attraction between ions

Intermolecular forces are the forces of attraction between molecules.
o There are three types of intermolecular forces: **dipole–dipole forces, London dispersion forces, hydrogen bonds,**

Dipole–dipole forces and London dispersion forces are known, collectively, as van der Waals forces.

o London dispersion forces exist between all molecules and are **relatively weak.**

e⁻ move

o Dipole–dipole forces exist only between **polar molecules.** They are **intermediate** in strength.

H-bond
dipole-dipole
dipole
dispersion

o Hydrogen bonds are dipole–dipole forces between an electronegative atom and a hydrogen atom
that is bonded to another electronegative atom on another molecule.
(Need bond between H and F, O, or N)

o Electrostatic forces are the forces that hold ions together within **ionic compounds.**

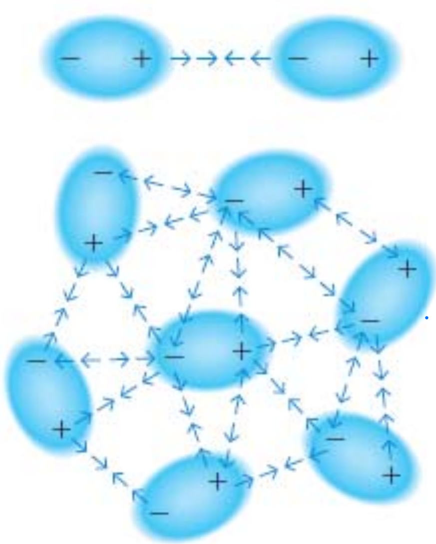


Figure 7 London dispersion forces are caused by a temporary imbalance in the position of electrons. These intermolecular forces can exist between single atoms or molecules.

As intermolecular forces increase, melting point and boiling point increase.

LEARNING TIP

The Strength of Bonds and Forces
Remember that intermolecular forces are very different from covalent bonds. Intermolecular forces act *between* molecules, attracting them toward each other. Covalent bonds hold the atoms *within* a molecule together. Even the strongest intermolecular force—the hydrogen bond—is much weaker than a covalent bond.

If we assign ionic and covalent bonds a strength of 100 on a relative scale, the intermolecular forces would all be below 10 on the scale.