Nanyang Girls' High School

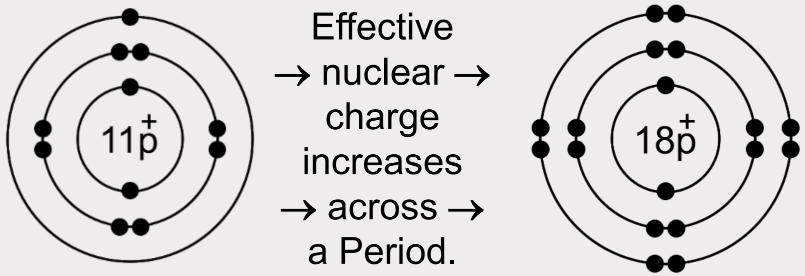
The Periodic Table of Chemical Elements
Trends Across Periods

Name: ....
Class: ....
Date: ....

1 **H** **hydrogen** **1.0**	• A *Period* is a horizontal row of elements in the Periodic Table. Elements in the same Period have the *same number of electron shells*. For example, the Period 3 elements sodium, magnesium, aluminium, silicon, phosphorus, sulfur, chlorine and argon all have three electron shells.	2 **He** **helium** **4.0**

• Elements are arranged in order of increasing atomic number.

• Across a Period, from left to right, the *proton number increases*. This means that the number of protons in the nucleus increases, so the *nuclear charge increases*. At the same time, electrons are added to the same outer electron shell (valence electron shell). Since the number of electron shells remains the same, the *shielding effect remains similar*.



• Because nuclear charge increases while shielding remains similar, the *effective nuclear charge experienced by the valence electrons increases*. As a result, the *electrostatic force of attraction between the positively charged nucleus and the negatively charged valence shell electrons becomes stronger*. This pulls the valence electrons closer towards the nucleus. Therefore, *atomic radius decreases across a Period*:

Na (11 p^+) = 186 pm, **Mg** (12 p^+) = 160 pm, **Al** (13 p^+) = 143 pm, **Si** (14 p^+) = 118 pm, **P** (15 p^+) = 110 pm, **S** (16 p^+) = 102 pm, **Cl** (17 p^+) = 99 pm

Note: The atomic radius for **Ar** (18 p^+) = 71 pm, although values for the *van der Waals radius* (188 pm) are sometimes given. 1 pm = 1×10^{-12} m.

• This stronger electrostatic force of attraction also affects metallic and non-metallic character. Metals tend to *lose electrons* to form positively charged ions, while non-metals tend to *gain electrons* to form negatively charged ions. Across a Period, the stronger electrostatic force of attraction between the nucleus and valence electrons means that atoms *lose electrons less easily* / *gain electrons more easily*. Therefore, *metallic character decreases / non-metallic character increases* from left to right across a Period.

• Across a Period, the nature of oxides also changes. Metallic oxides are usually *basic*, while non-metallic oxides are usually *acidic*. Therefore, across Period 3, the oxides change from basic to amphoteric to acidic: Na_2O and MgO are basic, Al_2O_3 is amphoteric, SiO_2, P_4O_{10} and Cl_2O are acidic.

• Melting points increase from Na to Mg to Al, which have strong metallic bonds. Si has the highest melting point due to the strong covalent bonds throughout its giant covalent structure. P_4, S_8, and Cl_2 have the lowest melting points in Period 3 due to the weak intermolecular force of attraction between their simple covalent molecules. Note: the melting point of S_8 is slightly higher than P_4 and Cl_2 due to its higher relative molecular mass.

• Electrical conductivity is very good for the metallic elements, increasing from Na to Al due to the increasing number of electrons in the delocalised 'sea' of electrons. Si is a *semiconductor*, while S_8, P_4 and Cl_2 are *electrical insulators* with no delocalised / mobile electrons to carry charge.

Nanyang Girls' High School

The Periodic Table of Chemical Elements
Trends within Groups

Name: ....
Class: ....
Date: ....

1
H
hydrogen
1.0

• A *Group* is a vertical column of elements in the Periodic Table. Elements in the same Group have the *same number of valence electrons*. This is why elements in the same Group have *similar chemical properties*. Down a Group, the proton number increases. This means that the number of protons in the nucleus increases, so the *nuclear charge increases*.

2
He
helium
4.0

• **Group 18**: Unreactive gases with fully filled valence electron shells.

• However, each element down the Group has *one more electron shell* than the element above it. Therefore, the valence electrons are found further away from the nucleus. The increase in the number of inner electron shells causes the *shielding effect to increase*. Inner electron shells shield the valence electrons from the full positive charge of the nucleus.

• **The Transition Metals**: Examples include manganese, Mn, iron, Fe, and Copper, Cu. The transition metals are characterised by their coloured compounds (blue $CuSO_4$), variable oxidation states (Fe^{2+} and Fe^{3+}) and catalytic properties (iron in the manufacture of NH_3).

• Although nuclear charge increases down a Group, the increase in shielding effect and the greater distance between the nucleus and the valence electrons reduce the attraction between them. Therefore, the *electrostatic force of attraction between the positively charged nucleus and the negatively charged valence electrons becomes weaker*. As a result, *atomic radius increases down a Group*. This affects the reactivity of Group 1 and Group 17 elements differently because Group 1 metals react by *losing electrons*, while Group 17 halogens react by *gaining electrons*.

• **Group 1 – The Alkali Metals**: Group 1 metals, *e.g.* sodium and potassium, have a single electron in their valence electron shell. They react by losing the single valence electron to form stable positively charged ions, *e.g.* $Na \rightarrow Na^+ + e^-$. Down Group 1, atomic radius increases and the shielding effect increases. The valence electron is further from the nucleus and electrostatic forces of attraction are weaker. Therefore, *less energy is required to remove the valence electron*. This means the atom loses its valence electron more easily. Hence, *reactivity increases down Group 1*. Group 1 metals react vigorously with water to form a soluble metal hydroxide and hydrogen gas, *e.g.* $2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$. The Group 1 metals are less dense than water and float on its surface. $H_2(g)$ from Na burns with a yellow flame, and $H_2(g)$ from K burns with a lilac flame.

• **Group 17 – The Halogens**: Melting points increases down Group 17 due to increasing relative molecular mass – Cl_2 is a yellow-green gas, Br_2 is a red-brown liquid and I_2 is a grey-black solid. Group 17 non-metals, *e.g.* chlorine and bromine, have seven electrons in their valence electron shell. They react by gaining one electron to form stable negatively charged ions, *e.g.* $Cl + e^- \rightarrow Cl^-$. Down Group 17, atomic radius increases and shielding effect increases. The incoming electron is added to a shell that is further away from the nucleus and more shielded by inner electron shells. Therefore, the electrostatic force of attraction between the nucleus and the incoming electron becomes weaker. This means *halogens gain electrons less easily down the Group*. Hence, *reactivity decreases down Group 17*. A more reactive halogen can displace a less reactive halogen from its halide solution. For example, chlorine can displace bromide ions: $Cl_2(aq) + 2NaBr(aq) \rightarrow Br_2(aq) + 2NaCl(aq)$ or $Cl_2(aq) + 2Br^-(aq) \rightarrow Br_2(aq) + 2Cl^-(aq)$.