Chem1311Ch10Ep1 Transcript

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Hello and welcome to the first episode of molecular geometry and orbital hybridization.

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The valence shell electron pair repulsion model.

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That's a mouthful.

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It is the best explanation for the geometries observed in molecules and its principles allow one to deduce the molecular geometry from a Lewis structure.

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VSEPR states that because electron pairs repel each other, chemical bonds and lone electrons alike will be located as far from each other as possible on the central atom surface.

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How far away electron pairs can be from each other depends on how many of them the central atom has.

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These electron pairs include both bonding electrons and lone pairs.

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We will be looking at the central atom as a sphere, upon which the electrons will sit.

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For simplicity, we will first consider atoms with only bonding pairs of electrons.

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We will include incomplete and extended octets as well.

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If a central atom has two pairs of bonding electrons, they will be furthest apart when they are.

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When they occupy opposite sides of the atom.

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As a result, the shape of the molecule will be linear.

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This is our first molecular shape.

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Beryllium chloride is an example of this molecular shape.

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Next we will consider 3 bonding pairs on the central atom.

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These electrons will be furthest apart from each other if they are 120 degrees from each other and will lie on the same plane.

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There is no possible way they can be on a different plane, because three points define a plane.

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This is the only way.

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The resulting molecular geometry is called trigonal planar.

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Because all atoms are bisected by the same plane.

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Here is the Equatorial view of this arrangement.

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And this is where the trigonal name comes from.

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So let's add it to our list.

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Boron trifluoride is an example of this type of molecule.

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Next, we will consider a central atom with four bonding electron pairs.

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In this case, placing the electrons at 90 degrees to each other will not maximize the distance between them.

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The distance between the atoms.

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Can be increased if they don't all lie on the same plane.

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Instead of a 90 degree angle.

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The angle can be increased to almost 110.

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By having the atoms lie in two separate planes.

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These planes will be perpendicular to each other.

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You can visualize this arrangement if you picture or maybe actually build 2 isosceles triangles with an obtuse angle of 110 and two acute angles of 35.

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Degrees each, both the same size.

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And place the obtuse angles facing each other like this slide shows.

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Also picture the atoms being at the acute angles with the central atom being at both obtuse angles.

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And then here's the important part.

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Rotate one of the triangles 90 degrees.

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And that's the shape.

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Let's add it to our list.

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This molecular shape is called tetrahedral.

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Methane is an example of a tetrahedral structure.

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Next, let's consider a central atom with five pair of bonding electrons.

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This is the polar view.

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We have two of the.

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Bonding electron pairs on the poles.

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And 3, 120 degrees away from each other at the equator.

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This is the Equatorial view.

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This shape is called trigonal bipyramidal.

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It is for two pyramids with triangular bases.

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Would look like if they were placed base to base.

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I've heard this called a diamond shape, but I have never seen a diamond.

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Shaped like this, actually.

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Let's add it to the list anyway.

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Phosphorus pentachloride is an example of this molecular shape.

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Finally, we will consider the molecular shape of an atom having six pair of bonding electrons on the central atom.

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They will be above and below front and back and left and right.

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Forming 90 degree angles all the way around.

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This shape is called an octahedral because it has eight sides.

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It looks like 2 pyramids with a square base connected base to base.

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Let's add it to the list as well.

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These are the five basic arrangements for electron pairs around the central atom.

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All the molecular geometries will be derived from these five basic shapes and will be caused by some of the.

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Bonding electrons not being bonding electrons.

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But rather lone pairs.

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Sulfur hexafluoride is an example of a molecule with an octahedral molecular shape.

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We will now look at the molecular shapes that result from having lone pairs on the central atom.

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This is the trigonal planar arrangement and.

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We will now consider.

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The shape of a molecule.

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Where one of these three electron pairs is a lone pair.

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The basic geometry of the electron pairs is unchanged.

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But the molecular shape is bent, and sulfur dioxide is an example of this shape.

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Next, we will consider a tetrahedral arrangement in which one of the electron pairs is a lone pair.

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The resulting shape is called trigonal pyramidal.

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Let's add it to the list.

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Ammonia is an example of this molecular shape.

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Next, we will consider a tetrahedral arrangement.

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In which two electron pairs are lone pairs?

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The resulting molecular shape.

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Is bent.

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Let's add it to our list.

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An example of this molecular geometry is water.

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Something to keep in mind is that lone pairs exert a greater repulsion than bonding pairs.

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So the angles between the bonding pairs of trigonal pyramidal and bent molecules are smaller than the angles of a tetrahedral molecule.

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Next, we will consider a trigonal bipyramid arrangement, in which one electron pair on the central atom is a lone pair.

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The resulting molecular shape is called a distorted tetrahedron.

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Here's another view.

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Let's add it to that growing list.

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Sulfur tetrafluoride is an example of this molecular shape.

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Next, let's consider another trigonal bipyramidal.

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Electron pair arrangement.

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In which two electron pairs are lone pairs?

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This molecular shape is called T-shaped?

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Let's add that to our growing list.

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Chlorine trifluoride is an example of this shape.

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Next, let's consider another trigonal bipyramidal shape.

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But this time, let's say that three of its electron pairs are lone pairs.

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This leaves us with a linear molecule.

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Let's add it to the list.

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Tri Iodide is an example of this shape of molecule.

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Next, let's consider a central atom of a molecule having an octahedral arrangement of electron pairs.

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If one of these electron pairs is a lone pair.

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The remaining atoms will form a square pyramidal molecular shape.

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Let's add this one to our list.

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Bromine pentafluoride.

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Is an example of this molecular shape.

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And finally.

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Let's consider an octahedral arrangement of electron pairs around the central atom, in which two pair.

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Are lone pairs.

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The resulting structure is called square planar.

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Let's add it to our list.

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Xenon pentafluoride

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I'm sorry.

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Xenon tetrafluoride is an example of this molecular geometry arrangement.

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And that's all there is.

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There isn't any more.