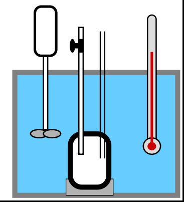
ENERGY FROM HYDROCARBONS

Hydrocarbons and Heat

- Most hydrocarbons are used as fuels.
- knowing how much energy a fuel provides, can tell us if it is useful for a certain application.
- For example, the amount of energy a food releases when burned, can tell us about it's caloric content (fats release lots of energy).
- Heat energy released during combustion can be measured with a calorimeter.
- A "bomb calorimeter" is shown. It includes water in a heavily insulated container, a stirrer, valve, bomb chamber, ignition wires, & a thermometer (pg 577).



Exothermic and Endothermic changes

- An alternative to the bomb calorimeter is a "coffee cup" calorimeter, where two nested polystyrene cups take the place of the container
- In either case, the change in heat of the water tells us about the reaction of the chemicals.
- An increase in water temperature indicates that the chemicals released energy when they reacted. This is called an "exothermic" reaction.
- In an "endothermic" reaction, water temperature decreases as the chemicals absorb energy.
- We will see that heat is measured in Joules (J) or kiloJoules (kJ). Before we do any heat calculations, you should know several terms …

<u>Specific heat capacity</u>
The heat needed to ↑ the temperature of 1 g of a substance by 1 °C. Symbol: c, units: J/(g°C).

<u>Heat capacity</u>
The heat needed to ↑ the temperature of an object by 1 °C. Symbol: C (=c x m), units: J/°C

<u>Heat of reaction</u>
The heat released during a *chemical reaction*. Symbol: none, units: J.

<u>Specific heat (of reaction)</u>
The heat released during a *chemical reaction* per gram of reactant. Symbol: h, units: J/g.

<u>Molar heat of reaction</u>
The heat released during a *chemical reaction* per mole of reactant. Symbol: ΔH, units: J/mol.

Heat Calculations

- To determine the amount of heat a substance produces or absorbs we often use $q = cm\Delta T$
- q: heat in J, c: specific heat capacity in J/(g°C), m: mass in g, ΔT: temperature change in °C,
- This equation makes sense if you consider units

$$J = \frac{J}{\cancel{g}\cancel{^\circ C}} \times \cancel{g} \times \cancel{^\circ C}$$

For a list of c values see page 568 (table 3) balloon demo

<u>Sample problem:</u>

When 12 g of a food was burned in a calorimeter, the 100 mL of water in the calorimeter changed from 20°C to 33°C. Calculate the heat released.

More practice

1. 5 g of copper was heated from 20°C to 80°C. How much energy was used to heat the Cu?
2. If a 3.1 g ring is heated using 10.0 J, it's temp. rises by 17.9°C. Calculate the specific heat capacity of the ring. Is the ring pure gold?
3. Do questions 5, 6 on pg. 596
   Do questions 7, 8 on pg. 570

Heat Capacity Calculations

- Recall that heat capacity (J/°C) is different from specific heat capacity (J/g°C).
- Heat capacity is sometimes a more useful value
- For example, because a calorimeter includes wires, the stirrer, thermometer, etc. some heat will be transferred to these other materials.
- Rather than having to calculate q for each material (like question 8) a J/°C value is used.

<u>Sample problem:</u>

A calorimeter has a heat capacity of 2.05 kJ/°C. How much heat is released if the temperature change in the calorimeter is 11.6°C?

Thermochemical Equations

Thermochemical equations are chemical equations with an added heat term.

- $KBrO_3(s) + 42\ kJ \rightarrow KBrO_3(aq)$
  This is endothermic (heat is absorbed/used)
- $2\ Mg(s) + O_2(g) \rightarrow 2\ MgO(s) + 1200\ kJ$
  This is exothermic (heat is produced/released)

Read 12.3 (pages 582 – 585). Do 2-6 (pg. 585).

<u>Sample problem:</u>

3.00 g of octane was burned in a calorimeter with excess oxygen, the 1000 mL of water in the calorimeter rose from 23.0°C to 57.6°C. Write the thermochemical equation for octane, representing the molar heat of combustion.